DOCUMENT RESUME

ED 260 930 SE 045 970

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TITLE Chemistry. Student Investigations and Readings. Investigations in Natural Science.
INSTITUTION Costa Rica Univ., San Jose. Research Inst. for the Improvement of Costa Rican Education.; Norman Independent School District 29, Okla.
PUB DATE 85
NOTE 284p.; For teacher's guide see SE 045 969. A joint project of the Science Education Center, University of Oklahoma and the Science Department, Senior High School of Norman Public Schools.
PUB TYPE Guides - Classroom Use - Materials (For Learner) (051)

EDRS PRICE MF01/PC12 Plus Postage.
DESCRIPTORS *Chemistry; High Schools; *Learning Processes; *Science Activities; Science Education; *Science Experiments; *Secondary School Science

ABSTRACT

Investigations in Natural Science is a program in secondary school biology, chemistry, and physics based upon the description of science as a quest for knowledge, not the knowledge itself. This student manual contains the 19 chemistry investigations. These investigations focus on concepts related to: interactions with water; salt and calcium; physical properties; heating and cooling a pure substance; composition of a compound; atomic theory and matter; lead nitrate and potassium iodide; combinations of zinc; hydrogen peroxide and manganese dioxide; mass relationships; temperature changes and dissolving; properties of ionic substances; Arrhenius solutions; pH, acids, and bases; acid-base interaction; salts and water; characteristics of methanol; solubility and structure; and periodicity and reactivity. Each investigation (which is accompanied by one or more readings with questions to answer) includes activities related to the three learning cycle components, namely, exploration, conceptual invention (the idea), and expansion of the idea. (JN)

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INVESTIGATIONS IN NATURAL SCIENCE

ED260930

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CHEMISTRY

Student Investigations and Readings

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INVESTIGATION 1: INTERACTIONS WITH WATER

Name_____

Section_____

INTRODUCTION

"Interaction" denotes the effect that two or more substances have on each other when they are combined. You have probably observed that oil floats on water, that dirt is carried along by water in rivers, and that fruit punch soft drink powder causes water to be a reddish-orange color. Although you will not be using oil, dirt and soft drink powder in the laboratory, some of the interactions that you will perform will resemble those interactions. In what ways are the interactions of oil, dirt and soft drink powder with water similar to each other? In what ways are they different? How would you try to convince someone that your stated similarities and differences were reasonable?

In this investigation you will combine water and one other substance in several separate experiments. Your tasks will be to determine the extent to which various substances interact with water and to determine some basic similarities and differences in the interactions of substances with water.

EXPLORATION

1. Put approximately 100 ml of distilled water in a clean, dry 250 ml beaker. Dip one piece of blue litmus paper, then one piece of red litmus paper in the water and record your observations.

2. Place the conductivity apparatus in the beaker, according to your teacher's instructions. Plug in the apparatus and record your observations.

   NOTE: Save the distilled water to use in the remainder of the tests.

3. Your teacher will provide you with various liquid and solid substances. For each test use 20 ml of water and 3 ml of liquid (or a pea-sized amount of solid) in a 50 ml beaker.

4. For each test describe three things:

   a. What happens to the substance immediately when it is placed in the 50 ml beaker with the water?

   b. What happens to the substance after it has been stirred vigorously?

   c. Describe what effect the mixture has on litmus paper and whether the mixture will conduct electricity.

5. Record your observations in the table below.

Substance (Name and Formula)	Before Stirring	After Stirring	Litmus/ Conductivity

6. When the tests are completed, dispose of the mixtures according to your teacher's instructions.

INVESTIGATION 1: INTERACTIONS WITH WATER

THE IDEA

1. Develop a classification system to describe various ways that substances interact with water.

2. What evidence can be used to show that some of these materials interact (mix) with water?

3. Considering the systems in which the water interacted with the other substance, make a statement describing the extent to which the substance mixes with the water in different regions of the mixture.

INVESTIGATION 1: INTERACTIONS WITH WATER

Name_____

Section_____

EXPANSION OF THE IDEA

EXPERIMENT

1. At each laboratory station there is a beaker containing a liquid. Perform the indicated tests to determine whether the liquid is a solution.

Substance	Color and Clarity	Litmus Paper	Conductivity	Solution (yes or not enough evidence)
1				
2				
3				
4				
5				
6				
7				
8				
9				
10				

INVESTIGATION 1: INTERACTIONS WITH WATER

EXPANSION OF THE IDEA

READING 1a: THE PROBLEM WITH SOLUTIONS

In Expansion of the Idea your investigation may have lead you to the conclusion that sucrose ($C_{12}H_{22}O_{11}$) and water did not form a solution. Wait a minute! Don't you drink sugar in your iced tea? Well, even if you don't, you are probably aware that sugar does dissolve in tea or water. In other words, your observations have not given you enough evidence.

This is the same problem you faced when you looked at the combination of magnesium oxide, iodine, and charcoal (among others) with water. Did these substances dissolve at all? (They certainly didn't totally dissolve!) Should we list them as soluble or insoluble? Let's look at some ways, other than direct visual observation, that we might use to answer these questions.

Recall, first of all, that distilled water alone was colorless, did not conduct electricity, and did not change the color of either blue or pink litmus paper. Consider next the fact that the dissolving process causes changes in the substance that is being dissolved. The substance is at least being broken apart into particles small enough to be invisible. In some cases--potassium chloride for example--the solution has properties that water alone did not have.

It is exactly these changes in properties that can provide evidence for partial dissolving. As an example, look at the experiment performed with iron (III) chloride. If you happened to use too large a piece of iron (III) chloride, you noticed that not all of it dissolved. A small piece would, however, totally dissolve in the 20 ml of water. You see that the experiment could lead to different conclusions. By analyzing the properties of the iron (III) chloride system in which all the solid dissolves, we can have information (results) that can be compared to the liquid part of a system where dissolving is questionable. Recall that the liquid took on a yellow-orange color, conducted electricity, and turned litmus paper from blue to pink. Three changes in properties are observed here, any one of which would indicate that some dissolving had occurred. Understanding the significance of this additional information gives only a partial solution to our problem with solutions, however. A question still remains as to how we decide whether to call a substance soluble or insoluble.

There are several ways that this problem can be handled satisfactorily, depending on the objectives you have in mind. You may, for example, want to use property change criteria (as above), or combine property changes with visual observation. This method could yield results such as the following: iron (III) chloride is soluble, magnesium oxide and iodine are slightly soluble, and charcoal is insoluble. (This information should not be interpreted as being right answers for the investigation. In fact, some students would probably take issue with these designations, after reviewing their observations.) An important point to see here is that solubility designations are somewhat arbitrary.

Another way to approach the problem of solutions would be to obtain quantitative information. This means that you would find out how much substance dissolved in a certain amount of water. Reference tables of solubilities will often list the extent of dissolving by stating the number of grams of a substance that will dissolve in 100 ml of water. Decisions about soluble or insoluble designations could then be based on a specific number of grams of substance being dissolved.

INVESTIGATION 1: INTERACTIONS WITH WATER

READING 1a: QUESTIONS

1. What characteristics of solubility have we used to identify a solution?

2. If the solubility of potassium chloride is 35 g per 100 ml of water, how many grams should you be able to dissolve using 5 ml of water?

3. The sucrose ($C_{12}H_{22}O_{11}$) solution and distilled water had no effect on litmus paper, did not conduct, or change color. If you were given a clear, colorless liquid, explain how you could determine if it was water or a sucrose solution. NOTE: Tasting is not an option.

INVESTIGATION 1: INTERACTIONS WITH WATER

EXPANSION OF THE IDEA

READING 1b: HOW HOMOGENIZED IS HOMOGENEOUS?

Consider a solution of potassium chloride in water. Now since this is a solution, it must be, by definition, a homogeneous mixture of the KCl and water. A further refinement of this definition would include a statement about the size of the particles of solute within the solution mixture--they are of molecular or atomic size. It's no shocking matter to extend the solution definition in this way because we can be certain that the solute particles are at least small enough to be invisible, and we can expect that their size definition is appropriate.

To get a better picture of the meaning of the term homogeneous, let us compare a potassium chloride-water solution with a mixture of powdered charcoal and water. You will recall that powdered charcoal was insoluble in water, and so you would not expect this mixture to follow the definition of a solution. The potassium chloride in water mixture, however, should certainly fit the solution definition. The following data were collected from 5 ml samples of a potassium chloride-water solution (Table 1) and a charcoal-water mixture (Table 2). In each case three samples were taken; one near the top of the mixture, one near the middle of the mixtures, and one near the bottom of the mixtures. These samples were poured into preweighed watch glasses (W.G.) and heated to dryness on a boiling water bath. When the samples were dry, the watch glasses with residue were weighed again and the weight of solid was computed.

Table 1

position	wt. of W.G.	wt. of W.G. + KCl	wt. of KCl
top	22.66 g	23.16 g	0.50 g
middle	23.36 g	23.86 g	0.50 g
bottom	22.30 g	22.80 g	0.50 g

Table 2

position	wt. of W.G.	wt. of W.G. + charcoal	wt. of charcoal
top	22.83 g	22.84 g	0.01 g
middle	22.88 g	22.89 g	0.01 g
bottom	23.37 g	23.47 g	0.10 g

(NOTE: The samples analyzed in Table 2 were taken one right after another, after the mixture had been stirred and then allowed to settle for five minutes.)

It is obvious, from the data in Table 1, that there was even distribution of the potassium chloride solute throughout the solution. It is equally clear that the charcoal was not evenly distributed. One is a solution; the other is not. One is homogeneously mixed; the other is not. The potassium chloride and water solution, in other words, has a uniform composition throughout. Using the information in Table 1, we could determine that a liter of that solution would contain 100 g of KCl, 10 ml would contain 1 g of KCl, and so forth. No such information could be derived from the charcoal system, however.

We have seen that a common characteristic of solutions is that they are transparent. When particles of solute can be observed in the mixture, we have to assume (because of our solution definition) that either the substance is not soluble in the solvent or that there is not enough solvent present to hold all the solid in solution. Tests, such as the ones discussed in Reading 1a, would allow us to determine which situation applies. It should be remembered, however, that the solution--the part of mixture that satisfies our definition--will be clear (transparent). Notice that there is a difference between clear and

colorless. A solution can have a color, and that color will depend on the substance dissolved, but it will still be clear, if it is a solution. You have observed some examples of colored solutions already, including the iron (III) chloride solution.

Suspension is the term applied to a mixture such as that of charcoal and water. Compared to the small, molecular-sized particles of a solution the particles in a suspension are huge, sometimes thousands or even millions of times larger. The solvent is unable to break the solute particles into small enough pieces so that it can hold those pieces in the solvent. The heavy pieces therefore, are observed settling out of the solvent. Can this settling effect explain the weight data in Table 2?

There is one other type of mixture that should be noted here, and that is one called a colloidal dispersion. Colloidal dispersions are mixtures resembling solutions, in some cases, but containing solute particles that are intermediate in size between the extremely small particles of a solution and the very large particles of a suspension. In colloidal dispersions the colloids (small particles) are dispersed (spread out) in the solvent. Although the solvent is not able to break these particles down enough to form a solution, it can hold the colloid particles so that they don't settle out. Examples of colloidal dispersions include the mixture of fat with water in homogenized milk, the mixture of tiny clay particles in the water of some rivers and lakes, and gelatin mixed with water.

INVESTIGATION 1: INTERACTIONS WITH WATER

Name_____

Section_____

READING 1b: QUESTIONS

1. A 15 ml sample of a solution contains 3.0 grams of solute. How much solute would be dissolved in: (Show your work.)

   a. 1 liter

   b. 100 ml

   c. 1 ml

2. Explain how air can be considered as a solution.

3. Explain how air can be considered as a colloidal dispersion (include at least two examples).

INVESTIGATION 2: SALT AND CALCIUM

Name_____

Section_____

INTRODUCTION

This investigation will involve the interaction of two more substances with water. One major difference between this and Investigation 1 is that you will collect weight data on the solid substances before adding water and at the end of the exploration procedure. How would you expect the weight of a solid substance to change if water were added to it to make a solution and then the water were evaporated to leave a solid substance?

EXPLORATION

1. Clean and dry your 100 ml and 150 ml beakers and two watch glasses. Weigh each of them three (3) times. Record the weights to two (2) decimal places.

100 ml beaker and watch glass	150 ml beaker and watch glass
_____	_____
_____	_____
_____	_____
average_____	average_____

NOTE: Round your average weight to two decimal places. Always keep the same beaker and watch glass combinations.

2. In one of the beakers place a spoonful of salt (sodium chloride—NaCl), and in the other beaker place a designated amount of calcium metal (Ca).

INVESTIGATION 2: SALT AND CALCIUM

3. Weigh each beaker, watch glass, and its contents three times and record the weights to two decimal places.

100 ml system and NaCl	150 ml system and Ca
_____ g	_____ g
_____ g	_____ g
_____ g	_____ g
average_____ g	average_____ g

4. From your data, determine the weights of salt and calcium contained in the system. Show your calculations.

Salt

Calcium

5. Add 25 ml of distilled water to the beaker containing salt. Replace the watch glass and record careful observations of the interaction between salt and water.

6. After the observations are written, test the remaining solution for conductivity and with litmus paper. Record the results of these tests.

7. Transfer the salt solution to a ring stand. Heat the system strongly until it is completely dry. While this system is heating, you may begin working with the calcium system.

8. Add 25 ml of distilled water to the beaker containing calcium. Replace the watch glass and record careful observations of the interaction between calcium and water.

9. Test the remaining liquid for conductivity and with litmus paper. Record the results of the tests.

10. When the salt system is completely dry, remove it from the ring stand and place the calcium system on the ring stand. Heat the calcium system strongly until it is completely dry.

11. After the dried systems cool to room temperature, weigh each system and record the weights.

salt system	calcium system
_____ g	_____ g
_____ g	_____ g
_____ g	_____ g
average _____ g	average _____ g

12. Determine the weights of the substances remaining in the two beakers. Show your calculations.

INVESTIGATION 2: SALT AND CALCIUM

13. Calculate the percentage weight change in the salt system. Clearly label the values used in your calculation.

14. Calculate the percentage weight change in the calcium system. Clearly label the values used in your calculation.

THE IDEA

1. What differences are there between the calcium system and the salt system relating to the interaction with water?

2. What conclusions can be drawn about the interactions in these two systems?

INVESTIGATION 2: SALT AND CALCIUM

Name_____

Section_____

EXPANSION OF THE IDEA

DEMONSTRATION: PHYSICAL OR CHEMICAL?

A. Copper (II) sulfate and ammonium hydroxide

1. Describe the procedure.

2. Record your observations.

3. Is it a chemical or physical change? What did you observe to support that it was a physical or chemical change?

B. Aluminum and hydrochloric acid

1. Describe the procedure.

2. Record your observations.

3. Is this a physical or chemical change? What did you observe to support that it was a physical or chemical change?

INVESTIGATION 2: SALT AND CALCIUM

EXPANSION OF THE IDEA

READING 2: FOOD FOR THOUGHT AND BREAKFAST

Many common systems, which seem to be simple and uninteresting, can actually be very complex and exciting if they are carefully observed. Take, for example, a pan of water containing a raw, uncracked egg. That system would seem to be sufficiently uninteresting at first glance and it would probably continue to be unless we did something to it. We will, therefore, put our simple, uninteresting system on the stove and begin heating it.

"How complex and exciting!", you say, as the system begins to get warm. Well, maybe not yet, but we have only begun to observe. One of the first things we see is the formation of many small bubbles on the bottom and sides of the pan, as well as on the surface of the egg. After awhile, these bubbles rise to the surface and are gone from the system. We can also notice that the egg moves around in the pan; not very much at first, but then it moves more and more as the water gets hotter. Both of these observations are important, and explaining them would give us much useful information. We will, however, pass by those observations for now and center our attention on the "big" things that are happening to the water and the egg.

Before very long, during this heating process, there are sounds coming from the system. Some of these sounds are light, pinging sounds and others are louder, resembling a <u>very mild</u> explosion. What is causing the noise? Careful observation shows bubbles forming at the bottom of the pan, but these bubbles are behaving differently from the very first ones to form. We could watch the first bubbles come all the way to the surface, but most of the bubbles that are forming now disappear before they get to the surface! Since these bubbles are formed at the bottom (where the system is hottest) and collapse on the way to the surface (where the system is coolest), maybe the collapsing is caused by the water cooling the bubbles. If this happens rapidly, the collapsing of the bubbles (implosion) could cause a shock wave that we could hear. The bigger the collapsing bubble, the bigger the shock wave. As we continue to watch, we see that the bubbles get closer and closer to the surface before imploding. Finally, the water gets hot enough that it cannot cool the bubbles as they rise, and the bubbles break the surface of the water. This condition is, of course, called <u>boiling.</u>

INVESTIGATION 2: SALT AND CALCIUM

If you hold your hand at a safe distance above the pan of boiling water, you notice that your hand gets damp. If you hold something cold over the pan (a glass of ice water, for example), the glass gets very wet on the outside. It is no shock to find out that the liquid on the glass is water--maybe "flat" tasting, but still water.

Having looked carefully at the water in the system, let us now turn our attention to the egg. In discussing the egg, we will have to use information from other work and observations taken at other times. (This is usually necessary any time we make intelligent observations or do sensible experiments.) We know what the raw egg looked like inside before we started heating the system. The "white" was not white at all, but transparent, and both the yolk and white were fluid. What the inside will look like after the heating process will depend on how long we heat the egg, but we know basically that the "white" will be white and less fluid, and the yolk will be more yellow than it was and less fluid.

Let's assume, for the sake of discussion, that we stop the heating process several times and examine the inside of the egg. In reality, of course, this would require several eggs because we would have to break them to observe the inside. If the egg is examined soon after the heating process is begun, we will see little, if any, change from the "raw" state. Later examination, however, would show no change in the yolk, but the white would have begun to change from clear to white. Most of the observable change would be near the shell. As we continue to observe eggs later in the heating process, we notice that the change extends continually farther in toward the center. The changes occurring in the egg, as in the water, must be caused by the heat applied to the system. Furthermore, the application of heat seems to be fairly uniform on all parts of the egg; that is, the bottom part is not changing more rapidly than the top part. Fortunately, this egg heating process is predictable. You know, for example, that if you want a soft-cooked egg it will have to cook for about three minutes after the water begins to boil. If, on the other hand, you want a hard-cooked egg for breakfast, you can begin heating the system and go jog a mile or two.

We have seen clearly defined changes occurring in both of the major parts of our system. The changes, though caused by the same heating, have been quite different. The changes in the water will be called "physical" changes, and the changes in the egg will be called "chemical" changes. Let's note some differences between the two types of changes.

The process of boiling could be reversed by simply removing heat, but the egg did not become raw as a result of cooling down. The changes in the egg, therefore, were irreversible by physical means. There were changes in color when the egg cooked. Several other properties, such as texture, odor and taste, changed as a result of cooking the egg. (If you don't believe that the taste changed, you can perform the obvious experiment.) So, in summary, we see that the water that underwent change was still water after the change, but the egg (though still recognizable as the day's work of a hen) had changed in its chemical makeup. You have seen other observational evidence of chemical change, which you will recall by reviewing your laboratory experience.

It is important to note that one cannot always determine that a chemical change has occurred by observing only one evidence for chemical change. For example, a gas was produced when water boiled, but no chemical change occurred. It would also be possible to change the color of a metal by heating it without changing the metal chemically. It will be wise, therefore, to know conditions under which changes take place before deciding whether the change is physical or chemical. It will also be wise to look for two or more evidences of chemical change before making your decision.

Although we have looked rather closely at this system, there are still many questions at which we have not even looked. For example:

- Why were there waves in the water when it was being heated?
- Why does the water that formed on the cold glass taste flat?
- What is the source of the stream of bubbles coming from one end of the egg?
- How does the egg get heated so uniformly?

As we continue our study of chemistry, it is quite possible that we will be able to give reasonable answers to these and other questions that could be asked about our complex and exciting system. Okay, so maybe it still isn't exciting, but at least it's complex (and informative).

INVESTIGATION 2: SALT AND CALCIUM

Name_____

Section_____

READING 2: QUESTIONS

For each of the following situations, state whether the process is chemical or physical, and then list supporting evidence.

1. Fermentation.

2. Leaves turning brown in the fall.

3. Running a gasoline engine.

4. Cooking a piece of bacon.

5. Popping popcorn.

6. Effervescent antacid tablet in water.

INVESTIGATION 3: PHYSICAL PROPERTIES

Name_____

Section_____

INTRODUCTION

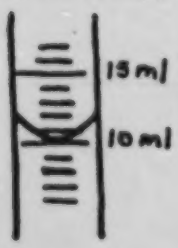
In this investigation you will measure the weights and volumes of three metals and one liquid. The volumes of the solids will be obtained by putting them in water and measuring the volume of water displaced (or pushed out of the way). This is a convenient method for determining the volume of an irregularly shaped object. Why is it that the water displacement method cannot be used to find the volume of some substances?

EXPLORATION

1. Obtain samples of lead (Pb), zinc (Zn), and aluminum (Al). Compute the average weight of each sample and record the weights in the appropriate spaces below.

	Lead	Zinc	Aluminum
	_____	_____	_____
	_____	_____	_____
	_____	_____	_____
average	_____	_____	_____

2. Determine the volume of each sample separately as follows:



a. Fill a 25 ml graduated cylinder to the 10 ml mark. Liquid levels should be read from the bottom of the dip (or meniscus)--see the diagram at left.

b. Carefully add one of the samples of metal to the graduated cylinder. Be sure that no air bubbles remain trapped on the metal.

c. Find the sample volume by subtracting the initial liquid level from the final liquid level:

(water and sample in ml) - (water alone in ml) = (sample in ml)

d. Repeat this procedure for all three samples and record the sample volumes in the spaces below.

lead_____ml

zinc_____ml

aluminum_____ml

e. Put the used metal samples in the place designated by your teacher.

3. Clean and dry a 10 ml graduated cylinder. Weigh the graduated cylinder accurately. Compute its average weight and record the weights in the appropriate spaces below.

_____

_____

_____

average _____

4. Your teacher will provide you with a sample of tert-butyl alcohol, which should be poured into your 10 ml graduated cylinder. Read the volume of this sample carefully and record it below.

tert-butyl alcohol_____ml

5. Now weigh the graduated cylinder containing the liquid sample. Compute its average weight and record the weights in the appropriate spaces below.

_____

_____

_____

average _____

6. Subtract the weight of the graduated cylinder (Step 3) from the weight obtained in Step 5 to obtain the weight of the liquid sample. Record that calculated weight below.

tert-butyl alcohol _____ g

7. Transfer your sample of tert-butyl alcohol to a clean, dry test tube. Stopper the test tube, label it, and put it away for use in the next laboratory session.

8. In order to utilize the data you have gathered in a productive way, it will be necessary to operate with your data, as well as the data from other students in your class. Begin by listing your data in the spaces provided below and in the table of Step 9.

NOTE: The third blank under each substance should contain a value obtained by dividing the weight of a substance by the volume it occupied.

	g	ml	g/ml
lead	_____	_____	_____
zinc	_____	_____	_____
aluminum	_____	_____	_____
t-butyl alcohol	_____	_____	_____

INVESTIGATION 3: PHYSICAL PROPERTIES

9. For purposes of comparison and analysis, list the data collected by the other students in your class.

lead		zinc		aluminum		t-butyl alcohol	
g	ml	g	ml	g	ml	g	ml

INVESTIGATION 3: PHYSICAL PROPERTIES

THE IDEA

1. What relationship exists between the weight and volume data for each substance?

2. Compute the <u>average</u> g/ml values for each of the four substances.

lead _____ g/ml

zinc _____ g/ml

aluminum _____ g/ml

t-butyl alcohol _____ g/ml

3. What information can be gained from the g/ml values?

INVESTIGATION 3: PHYSICAL PROPERTIES

Name_____

Section_____

EXPANSION OF THE IDEA

DEMONSTRATION I: COMPARING DENSITIES

1. Describe the apparatus.

2.

	Substance A	Substance B
Weight of graduated cylinder	_____	_____
Weight of graduated cylinder and liquid	_____	_____
Weight of liquid	_____	_____
Volume of liquid	_____	_____

3. Determine the density of each of the substances. Show your work.

INVESTIGATION 3: PHYSICAL PROPERTIES

QUESTIONS

1. Are Substance A and Substance B the same substance? Justify your answer with experimental evidence.

2. What other tests might you use to determine if these liquids are the same substance?

INVESTIGATION 3: PHYSICAL PROPERTIES

Name_____

Section_____

EXPANSION OF THE IDEA

DEMONSTRATION II: COMPARING DENSITIES

1. Describe the contents of graduated cylinder #1.

2. Describe the contents of graduated cylinder #2.

QUESTIONS

1. Give a logical explanation for the appearance of the liquids in graduated cylinder #1.

2. Give a logical explanation for the appearance of the liquids in graduated cylinder #2.

INVESTIGATION 3: PHYSICAL PROPERTIES

EXPANSION OF THE IDEA

READING 3: HIERO'S HERO

It has probably puzzled everyone at one time or another that a ship made of steel and weighing several thousand tons could ever float. Of course, there is a very good reason for being puzzled. Metal objects typically do not float on water, and that is because their densities are greater than that of water. In fact, it is generally observed that all objects having a greater density than water will sink. (Exceptions would include objects that are soluble in or would react chemically with water.)

If the ship that we are considering were a solid block of metal, it would certainly sink! In fact, a ship does sink into the water partially. Imagine lifting a ship out of the water. If the water were in the solid state (ice), you would observe a hole in the ice that would give an exact outline of the part of the ship that had been under water. No such hole would be observed, however, if the ship were lifted out of liquid water because the water would fill in the space that had been occupied by the ship. We see, therefore, that for the ship to make the <u>hole</u> in the water, it must have pushed water out of the way. This is called <u>displacing</u> water.

A valid question to ask at this point is, "Why does the ship only sink into the water partially and then stop sinking?" An interesting result is obtained when you consider the amount of water that is displaced by the ship. If, for example, we could measure the volume of our hypothetical hole in the ice and multiply that by the density of the water, we would know the weight of water displaced (NOTICE: ml x g/ml = g). Upon doing this calculation, we arrive at our interesting result: The weight of water displaced by the ship is <u>equal</u> to the weight of the ship!

If you have ever watched someone get into a canoe, you have seen an illustration of this principle. The added weight of the person getting into the canoe causes the total weight of the system to increase. As a result, the canoe sinks farther down into the water until a weight of water equal to the system weight has been displaced.

Now let's return to our ship example again and consider the validity of a statement made earlier about the shape of the ship. Assume, for simplicity, that our ship is made only of iron and that this iron weighs 7,870,000 kg (which is 7.87×10^6 kg or 7.87×10^9 g). If we wanted to

form this metal into a cube, we could determine the size of the cube by knowing the density of the iron (7.87 g/cc) and the weight of the metal. Follow the computation carefully.

$$\frac{7.87 \times 10^9 \text{ g}}{7.87 \text{ g/cc}} = 1 \times 10^9 \text{ cc}$$

This result is the <u>volume</u> (space) occupied by the metal in our ship (one billion cubic centimeters). Further, assume that we are going to try to float this block in pure water having a density of 1.0 g/cc. From our previous discussion we recall that this block will have to displace an amount of water equal to its own weight in order to float. Stated in numbers, we are saying

$$\frac{7.87 \times 10^9 \text{ g}}{1 \text{ g/cc}} = 7.87 \times 10^9 \text{ cc}$$

is the volume of water that must be displaced by the iron to make it float. When we look back at the total volume of iron we have to use in displacing water, it is immediately clear that our block <u>does not</u> have enough volume to displace the water necessary to allow it to float. The block would have to be almost eight times as big as it is (with no increase in weight) to float.

How much water <u>would</u> our cube of iron displace in sinking to the bottom? Well, if the iron doesn't absorb water (which it doesn't) and the cube sinks (which it does), then the volume of water displaced would have to equal the volume taken up by the cube (which it does). You have already used this principle of displacement to determine a <u>characteristic property</u> --density--for several substances.

A story is told of the supposed discovery and first use of this idea by Archimedes in about 250 B.C. The ruler of Syracuse, King Hiero II, had had a crown made from pure gold. He had some reason to suspect the honesty of the goldsmith who had made the crown, so he called in Archimedes--a respected thinker and experimenter--and ordered him to determine if the crown was made of pure gold or if the gold had been mixed with cheaper metals. What made this problem difficult to solve was that King Hiero did not want the crown to be damaged in any way.

INVESTIGATION 3: PHYSICAL PROPERTIES

Archimedes knew that equally-sized pieces of gold and other metals--copper and silver, for example--weigh different amounts (they have different densities). According to the story, the importance of the relationship between weight and volume in solving this problem became apparent to Archimedes, as he lowered himself into a bathtub and noticed the water level rising. His body had displaced water in the tub, as would any solid object. The realization of a method for solving his problem excited Archimedes so much that he jumped out of the bath and went streaking through the streets of Syracuse yelling "Eureka!" (meaning "I've found it", or "I have it"). When he settled down a little bit, he weighed the king's crown, put it in a container of water, and noted the amount of displacement. Then he obtained an equal weight of pure gold, put it in water, and observed that the displacement was <u>less!</u> This showed that the king's crown (being less dense than pure gold) had been made from a mixture of metals.

INVESTIGATION 3: PHYSICAL PROPERTIES

Name_____

Section_____

READING 3: QUESTIONS

SHOW ALL WORK, WHEN APPROPRIATE.

1. A solid brick, weighing 10 grams, has the following dimensions:

   Length = 2 cm; width = 1 cm; height = 1 cm.

   a. What is the density of the brick?

   b. If the brick were cut in half, what would be the density of each piece?

2. You are collecting a gas by displacement of water and the gas is slightly soluble in water.

   a. How does this affect the volume of gas you collect?

   b. How does this affect your determination of the density of the gas? Explain.

3. Draft readings on ships tell how much of the ship is under water. For purposes of this question, assume that the total weight of the ship remains the same and that the ship is not touching the ground. At one time a draft reading of 25 feet-7 inches is made, and at another time a draft reading of 27 feet-1 inch is made. Explain the differences in the draft readings. (HINT: The two readings cannot be taken at the same geographic location.)

INVESTIGATION 3: PHYSICAL PROPERTIES

Name_____

Section_____

EXPANSION OF THE IDEA

DENSITY PROBLEMS

Solve the following problems. Be sure to show <u>all</u> your work and include the correct units.

1. If a piece of clay has the measurements 2 cm x 4 cm x 6 cm and weighs 96 grams, calculate its density.

2. Nine liters of gasoline have a mass of 6120 grams. What is the density of the gasoline?

3. What is the weight in grams of a block of silver measuring 2.5 cm x 8.0 cm x 4.0 cm, given that the density of silver is 10.5 g/cm^3?

4. Calcium chloride has a density of 2.15 g/cm^3. What would be the volume of 3.37 g of this substance?

5. If 2 liters of milk weighs 2.06 kg, calculate its density.

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

Name_____

Section_____

INTRODUCTION

There are three phases (or states) of matter that we consider in this course. They are solid, liquid and gas. In this investigation you will be focusing your attention on changes between phases, such as liquid to solid and liquid to gas. Are these phase changes physical or chemical in nature? What evidence can you present to support your choice?

EXPLORATION

1. Pour 3 to 5 ml of your tert-butyl alcohol sample into a clean, dry 10 ml graduated cylinder. Dispose of the remaining part of your sample according to your teacher's instructions. Record the volume you will be using. _____ ml.

2. Return the measured portion to the test tube in which it was originally.

3. Fill a 400 ml beaker half full of tap water. Measure the temperature of the water and, if necessary, add ice to lower the temperature of the water to 15°C. When the temperature has reached 15°C, remove any remaining ice from the water.

4. Dry your thermometer and then insert it into the test tube containing the tert-butyl alcohol sample. Stir the sample with the thermometer until the temperature stops changing. Record this temperature on the Freezing Data table. (Time=0 minutes.)

5. When you have recorded your initial temperature, lower the test tube into the water bath, continue stirring the sample constantly, and record temperature readings every 15 seconds.

6. Continue taking temperature data until one minute after all the liquid has solidified. If the sample becomes hard to stir, do not force the thermometer in the solid to the point of breaking the thermometer.

7. Mark with an asterisk (*) the temperature at which solid first appears and the temperature at which the sample is totally solidified.

FREEZING DATA

time (min)	temp (°C)	time (min)	temp (°C)	time (min)	temp (°C)
0.00	____	2.25	____	4.50	____
0.25	____	2.50	____	4.75	____
0.50	____	2.75	____	5.00	____
0.75	____	3.00	____	5.25	____
1.00	____	3.25	____	5.50	____
1.25	____	3.50	____	5.75	____
1.50	____	3.75	____	6.00	____
1.75	____	4.00	____	6.25	____
2.00	____	4.25	____	6.50	____

8. Set up a ring stand with a large iron ring and a wire screen with an asbestos center. Place the 400 ml beaker containing water on the ring stand. Clamp the test tube containing the tert-butyl alcohol sample on the ring stand so that the test tube does not touch the sides or the bottom of the beaker.

9. Begin heating the system with a bunsen burner. When the solid melts, remove the thermometer from the test tube, dry it, and insert it in the water bath, being careful to keep it from sustained contact with the bottom and sides of the beaker. Put one boiling chip in the test tube containing the tert-butyl alcohol sample and one in the water bath.

10. When the water temperature reaches 70°C, remove the thermometer from the water bath and dry it. Continue heating the system.

11. Hold your thermometer just <u>above</u> the liquid in the test tube and record temperature readings (to the nearest half degree) every 15 seconds until one minute after all the liquid has boiled out of the test tube. Mark with an asterisk (*) the temperature at which the liquid begins boiling <u>and</u> the temperature at which all the liquid has boiled away.

<u>BOILING DATA</u>

time (min)	temp (°C)	time (min)	temp (°C)	time (min)	temp (°C)
0.00	____	2.75	____	5.50	____
0.25	____	3.00	____	5.75	____
0.50	____	3.25	____	6.00	____
0.75	____	3.50	____	6.25	____
1.00	____	3.75	____	6.50	____
1.25	____	4.00	____	6.75	____
1.50	____	4.25	____	7.00	____
1.75	____	4.50	____	7.25	____
2.00	____	4.75	____	7.50	____
2.25	____	5.00	____	7.75	____
2.50	____	5.25	____	8.00	____

12. Set up coordinate axes on graph paper to plot the data for freezing and boiling. Graph the two sets of data on separate graphs, with time on the horizontal axis and temperature on the vertical axis. Your teacher will assist you in getting started with the graphing.

13. When the points have been plotted, draw a smooth line through consecutive points on the graph. Add the asterisks above the appropriate points on the graph.

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

THE IDEA

1. You will observe that each of the graphs can be divided into three regions of temperature. Label the regions "a", "b" and "c" from left to right on each graph. What phase(s) of tert-butyl alcohol exist(s) in each region?

   freezing graph: a._____ b._____ c._____

   boiling graph: a._____ b._____ c._____

2. In the spaces below, record your sample volume and the temperatures corresponding to region "b" of the two graphs. List those data collected by other students in your class.

sample volume (ml)	freezing region "b" (°C)	boiling region "b" (°C)
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____
_____	_____	_____

3. What relationship exists between sample volume and the boiling temperatures listed?

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

Name_____

Section_____

EXPANSION OF THE IDEA

EXPERIMENT: IDENTIFYING AN UNKNOWN SUBSTANCE

1. You will be given a sample of a substance which you are to identify using melting points and/or boiling points. Your sample will be among those listed in the table below. Complete the table by filling in the right-hand column of the table.

SUBSTANCE	M.P. (°C)	B.P. (°C)	PHASE AT ROOM TEMPERATURE
para-dichlorobenzene	53	174	_____
phenol	43	182	_____
acetic acid	17	118	_____
cyclohexane	7	79	_____
acetamide	82	221	_____
n-propanol	-127	98	_____
naphthalene	80	211	_____
methanol	-97.8	64	_____

2. Use the basic freezing and boiling procedures that you used earlier. Record your data in the following table, then graph the data.

FREEZING/BOILING DATA

time (min)	temp (°C)	time (min)	temp (°C)	time (min)	temp (°C)

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

EXPANSION OF THE IDEA

READING 4: WIGGLES, VIBRATIONS, AND THE TEMPERATURE OF THINGS

Have you ever looked carefully at the vapor coming from a whistling tea kettle, as water is boiling in the kettle? If you have, you noticed two distinctly different regions in the gas coming from the hole in the kettle. Just above the hole the gas is transparent and invisible (because it is colorless). Farther away from the hole, however, the gas becomes foggy (translucent to opaque) and appears to be white.

If you held a thermometer in the white, foggy vapor until the temperature remained constant, and then moved the thermometer closer to the hole and waited to get a temperature reading, you would find that the temperature nearer the hole is higher than the first temperature taken. Assuming that the second reading was taken in the white fog also, you would find a third reading, taken in the transparent gas, to be even higher than the other two.

When you hold a glass rod in the white vapor, you notice that moisture collects on the rod. A different area of the rod held in the transparent gas also accumulates liquid. In both cases the liquid will be found to be water; the substance you started with in the tea kettle.

Let's summarize the entire system under discussion here. Energy is added to the water in the tea kettle. As heat is added the water boils (vaporizes, or turns to gas) and leaves the tea kettle. The farther away from the kettle the gas gets, the cooler it gets. A cold object (the glass rod) placed in the path of the gas causes the gas to change back to liquid water, whether the rod is placed in the warmer or the cooler region of the gas.

This system is quite common and familiar as far as we have taken it. Careful examination of the system and its changes, however, can illustrate some very useful ideas. We can get at these ideas by considering the following questions:

1. What happens to the particles of a liquid when the liquid boils?

2. What causes the particles of a gas to come together (or condense)?

3. How does the movement of molecules vary with temperature?

4. Could the movement of molecules be used to explain a temperature reading on a thermometer?

Let's deal with these questions one at a time, beginning with Question 1. We could begin our answer by simply saying that the particles of the liquid separate from each other. This can be seen in the vapor coming from the tea kettle. The farther it gets from the source, the more dispersed (or diffuse) it becomes. This separation of particles occurs as a result of a "disturbance" of the system. The energy added to the water (heat energy, in this case) allows the water molecules to overcome the attractions that are holding them together in the liquid state. When the attraction is overcome, the separated particles move rapidly and randomly.

The second question relates to the observations of water vapor condensing on the glass rod. Notice the similarity between this condensation and the condensation that occurs on the insides of windows during the winter, or on the outsides of cold drink glasses during the summer. In each case we have water vapor coming in contact with a relatively cold surface. Since the water had to gain energy to become vapor, it would stand to reason that the water would have to lose energy to become liquid again. The cold surface absorbs energy from the water vapor, and the attractive forces between the particles are able to hold the molecules in the liquid state.

The answer to Question 3 has been implied in the answers to the first two questions, but it needs to be expanded a bit further. If some small, lightweight objects (such as styrofoam beads) were placed in water, we would notice that these objects would begin moving faster and faster, as the temperature increased. We can assume that the individual water molecules (which we cannot see) move faster and collide harder and more often with the objects (which we can see) causing them to move faster. Someone might argue that heat is causing the objects to move around, not the movement of water molecules. We would have to agree, but only partly. If we were to place styrofoam beads in a beaker without water and begin heating the beaker, we would get a little motion and a lot of melted, smelly styrofoam. So the heat causes the objects to move around, but only because the heat energy has added to the energy of motion of the water molecules. This energy of motion (kinetic energy) is then transmitted to the objects through collisions. We see, therefore, that the amount of motion of molecules corresponds directly to the temperature of the substance. In other words, the higher the kinetic energy, the higher the temperature.

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

To pursue an answer to our fourth question, consider a thermometer placed in a container of liquid. Since the liquid particles are continuously moving and colliding with each other, those particles will also collide with the glass of the thermometer. These collisions will increase the kinetic energy of the glass particles which, in turn, collide with the particles of mercury and increase their kinetic energy. The mercury expands as a result of the increase in kinetic energy, and this expansion can be calibrated to read "temperature".

Let's examine the idea of expansion with heating, in terms of the model for matter which assumes that its particles are in constant motion. The slower an object is moving, the less space is needed to contain it. For example, the chemical reactions that occur in a fire cracker produce gases very rapidly. Since these gas particles are moving faster than the solid, unreacted particles, they exert a great force outward and break the barrier surrounding them. This activity shows up as an explosion. You have noticed that an explosion does not occur if you ignite the powder in a fire cracker that you have broken open. The expansion of the gases takes place much more freely in this latter situation.

As an analogy to the situation involving invisibly small particles, consider an auditorium filled with 500 people. If everyone is seated (or even standing), there is plenty of room for comfortable movement, such as adjusting body positions and minor stretching. Now allow the activity of the people to increase by having everyone do deep knee bends. This could be followed by having the people run in place. Clearly, these last two activities are not approprite for the auditorium, and they would surely result in bruised bodies and strained relationships among the people. Why are these activities inappropriate? Because, as movement (kinetic energy) increases, the space required for that movement also increases. Notice that the people do not change size, but the <u>effective space</u> that they occupy increases, as their activity increases.

Almost every substance expands as energy is added to it. Furthermore, this expansion is predictable, allowing the use of mercury (or some other substance) in a thermometer. Expansion of materials is also an important consideration in many types of construction, such as highways and bridges. Without expansion joints in these structures, they would buckle under the forces caused by expansion at higher temperatures.

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

We can conclude from all this discussion that all matter is in constant motion. This applies to solid, liquid, and gaseous states of matter. When we measure the temperature of a system, therefore, we are observing a measure of the <u>average kinetic energy</u> (energy of motion) of the particles within the system.

INVESTIGATION 4: HEATING AND COOLING A PURE SUBSTANCE

Name_____

Section_____

READING 4: QUESTIONS

1. If a glass of water is left uncovered at room temperature, it will eventually all evaporate. Given that the boiling point of water is 100°C, offer an explanation for the evaporation of the water.

2. What characteristic of a substance will determine what the value of its boiling point will be? Give information that will logically support your answer.

3. Explain why dew is found on grass in the early morning instead of in the early evening.

4. Explain how it is reasonable to assume that the particles within solid objects are in motion.

INVESTIGATION 5: COMPOSITION OF A COMPOUND

Name_____

Section_____

INTRODUCTION

This investigation will involve weighing at key points in the exploration. As usual, care taken with weighing the system and care exercised in not spilling the contents of the system are essential to getting valid data. You will also notice that you will be asked to use steel wool to scrape off the outer layer of your magnesium metal strip. How do you suppose the layer formed in the first place? How would leaving the film on the metal affect your experimental results?

EXPLORATION

1. Clean and dry your evaporating dish and crucible cover.

2. Accurately weigh the empty evaporating dish with the crucible cover to two decimal places.

_____

_____

_____

average _____

3. Obtain a strip of magnesium metal ribbon from your teacher. Scrape off the adhering film with fine steel wool and then wipe the ribbon with a clean paper towel. Cut the metal into small pieces and place them into the evaporating dish. Weigh the system to two decimal places.

_____

_____

_____

average _____

4. Using your data, calculate the weight of magnesium metal in the system. Show your work.

5. Heat the covered evaporating dish and magnesium system gently for two minutes, gradually increasing the heat intensity. Continue heating the system strongly for an additional 15 minutes, then turn off the burner.

6. As the system is cooling, visually inspect the contents of the system. Record your observations.

7. After the evaporating dish has cooled, add only enough distilled water to the dish to cover its contents. Gently heat the evaporating dish. As the water boils away, smell the vapors coming from the evaporating dish and identify them. After the water has evaporated, continue to strongly heat the system for five minutes. Heat to constant weight.

_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

8. Determine the weight change in the system. Show your work.

9. Record careful observations of the final contents of the evaporating dish.

10. You have now calculated two weights: the weight of magnesium and the weight gained (or lost) during reaction. There are four mathematical operations (addition, subtraction, division, multiplication) available to manipulate these weights. Use these operations on the weights and fill in the chart below. Include units, when appropriate.

wt. Mg + wt. chg.	wt. Mg - wt. chg.	$\frac{\text{wt. Mg}}{\text{wt. chg.}}$	wt. Mg x wt. chg.

INVESTIGATION 5: COMPOSITION OF A COMPOUND

THE IDEA

1. What substances could possibly have reacted with the magnesium? (that is, what are the substances available to the system?)

2. Which of the four mathematical operations in Step 10 would you expect to give the same value, no matter how much Mg was used? Why?

INVESTIGATION 5: COMPOSITION OF A COMPOUND

Name_____

Section_____

EXPANSION OF THE IDEA

EXPERIMENT: HEATING HYDRATED COPPER (II) SULFATE

1. Place a small amount of hydrated copper (II) sulfate in a test tube. Gently heat the test tube with a Bunsen burner. Record your observations.

2. Allow the test tube to cool and then add a few drops of water to the solid in the test tube. Record your observations.

3. Clean and dry your evaporating dish.

4. Accurately weigh the empty evaporating dish to two decimal places.

_____

_____

_____

average _____

5. Add between 1 and 4 grams of hydrated copper (II) sulfate to the evaporating dish and weigh the system to two decimal places.

_____

_____

_____

average _____

6. Calculate the weight of hydrated copper (II) sulfate added. Show your work.

7. Heat the evaporating dish gently for twelve to fifteen minutes, then turn off the burner.

8. As the system is cooling, visually inspect the contents of the system for evidence of complete reaction.

9. After the evaporating dish has cooled, weigh the system, and then heat gently to constant weight.

_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____
_____	_____	_____	_____

10. Determine the weight of the residue in your system. Show your work.

11. Determine the weight change in the system. Show your work.

INVESTIGATION 5: COMPOSITION OF A COMPOUND

QUESTIONS

1. The chemical formula for hydrated copper (II) sulfate can be written

   $CuSO_4(H_2O)_x$,

   where x represents a number of H_2O molecules. What are the possible products of heating this substance? Write equations for the various reactions predicted.

2. Did the weight increase or decrease when you heated the hydrated copper (II) sulfate:

   $CuSO_4(H_2O)_x$?

   Offer an explanation for the change.

3. Divide the weight change by the weight of the residue. Show your work and explain what this number represents.

4. Use the weights on the periodic table to determine the best value of "x" in

   $CuSO_4(H_2O)_x$.

   X should be an integer. Show your work.

INVESTIGATION 5: COMPOSITION OF A COMPOUND

EXPANSION OF THE IDEA

READING 5: THE CHEMICAL CONGRESSMEN

Congress for over two centuries has been instrumental in formulating the laws of the land. Those involved have gathered information, weighed facts, and have finally arrived at judgments affecting the nation as a whole. In a similar fashion, various chemists in centuries past have compiled laboratory observations, made hypotheses, developed theories, and established chemical laws. Just as civil laws may be stricken from the books because they are no longer applicable, even so chemical laws sometime must be revised when new experimental information is made available.

The history which surrounds the development of the foundations of modern chemistry is fascinating and no one man can be credited with all of the discoveries. Several scientists' lives and findings will serve to demonstrate three chemical laws that have been formulated.

The Congressman from France (1743-1794)

Although Antoine Laurent Lavoisier is viewed as the "father of modern chemistry", we should remember that he is not without predecessors who were instrumental in shaping his theories. Joseph Priestly (1733-1804), to whom we owe the fizz in carbonated beverages, had shown that when mercuric oxide (then called mercury calx) was heated, a gas was produced which caused a candle to burn "with a remarkably vigorous flame". He breathed some of this "air" and remarked after experiencing its invigorating effects, "Who can tell but that, in time, this pure air may become a fashionable article in luxury.*

Lavoisier repeated the experiment, naming the gas oxygen. Why should copying another chemist's experiment be so important to the scientific community? It shouldn't, except that our Frenchman examined another facet of the reaction that Priestly had not bothered to check. He weighed the system before and after the reaction. This may not sound like a momentous innovation (especially since we have been weighing nearly everything three times). In the eighteenth century, though, scientists were primarily concerned with qualitative descriptions of substances.

* A Short History of Chemistry, J. R. Partington, Harper and Brothers, New York, 1960

That is, they were interested in describing physical changes in systems, for example, the changing of a metal into a non-metal by the addition of heat. Visual observations are extremely important, but they are further enhanced by quantitative data. Isaac Newton and Galileo proved to physicists, initially, that measuring the size of objects, their mass, and the time it took them to move through a certain distance was important when analyzing the motion of the planets. Lavoisier, though not the first to do so, simply applied these principles to chemistry.

What did he find? Well, when the mercuric oxide was heated after accurate weighing, it lost weight. He had begun with 216 grams of a red powder and finished with 200 grams of quicksilver, liquid mercury. The gas was trapped and consequently weighed. . .you guessed it--16 grams of oxygen had been produced. The mass of the reactants equaled the mass of the products. We now call this principle <u>the law of the conservation of mass</u> in a chemical reaction.

216 g		200 g		16 g
mercuric oxide	-->	mercury metal	+	oxygen

Lavoisier went on to prove this law in reaction after reaction, but let us move on to the next chemist.

Yet Another Congressman from France (1754-1826)

Joseph Proust believed, on the basis of his experimental observations, that compounds were not haphazard in their composition. That is, a particular substance had particular portions of each element in it. (For example, mercuric sulfide contained a certain amount of mercury and a fixed amount of sulfur.) Again, this hypothesis may sound simplistic. But remember, chemistry was becoming quantitative at this time and many observations were still purely qualitative. Chemists and Alchemists alike knew well that mercury and sulfur were the components of mercuric sulfide, but some held that the amounts of each were widely flexible. Some of the confusion was brought on by the fact that metallic alloys and various glasses can be formed in indefinite proportions. This knowledge did not sway Proust. He contended rightly that these substances were mixtures (today we might also call them solid solutions) and not compounds produced by chemical reactions. How could he be so certain? Mr. Proust had based his strong convictions on many observations of chemical combination.

One such experiment dealt with copper (II) carbonate, $CuCO_3$. He obtained this compound from several natural sources and also prepared it in the laboratory. Each sample, regardless of its source, was found to contain a definite amount of copper combined with a definite amount of carbon and oxygen. Proust's research proved convincing to most of the scientists of his day and his theory is now known as <u>the law of constant composition.</u>

The Congressman from England (1766-1844)

John Dalton, by his own admission, was not a brilliant scientist. Perseverance in the laboratory was his primary asset, and we find that he had to be diligent considering his perceptual handicap. He was color blind, an obvious problem for the research chemist.

In the course of his study, Dalton analyzed many substances to determine their composition. He found that the proportion of oxygen to carbon in carbon dioxide--CO_2--was twice that in carbon monoxide--CO. Given two equal portions of carbon, he also discovered that twice as much methane gas--CH_4--could be produced as ethane gas--C_2H_4.

Various substances, then, not only had constant compositions as stated by Proust, but their elements were combined in whole number ratios. We call this principle <u>the law of multiple proportions.</u> It is interesting to note that although Dalton was the first to develop the law of multiple proportions, he was not able to apply it accurately to every substance he tested. He falsely reasoned that the most common combination of two elements in nature should be the simplest possible, one to one. On this assumption, water would have been known as HO. Ammonia, by the same premise, was NH instead of NH_3. Forty more years of research were necessary to clear up this problem of compound formulas, but Dalton proved well that <u>the law of multiple proportions</u> was here to stay.

Lavoisier, Proust, and Dalton compose an extremely small portion of all the "law makers" who have contributed to the body of chemical knowledge. These chemists simply set forth some of the first and most fundamental principles which aid us every time we enter the laboratory. More laws are left to be uncovered, more than have yet been revealed.

INVESTIGATION 5: COMPOSITION OF A COMPOUND

Name_____

Section_____

READING 5: QUESTIONS

1. Determine formulas for the following compounds of nitrogen and oxygen. Show your work.

Compound	wt. of nitrogen (g)	wt. of oxygen (g)
A	1	2.85
B	1	2.28
C	1	0.57
D	1	1.71

INVESTIGATION 5: COMPOSITION OF A COMPOUND

2. Samples of gas composed of carbon and oxygen were collected from two sources. Source A was exhaust gas from a gas burner, and Source B was an enclosed terrarium containing plants. Using the data below, determine the identity of the two gases (showing your work).

Source	wt. of carbon (g)	wt. of oxygen (g)
A	5.0	13.3
B	0.3	0.8

3. Which of the three laws discussed in this reading are illustrated by Question 1 and Question 2? Explain.

4. Which of the three laws discussed in this reading are illustrated by the heating of magnesium in Investigation 5? Explain.

5. Which of the three laws discussed in this reading are illustrated by the heating of hydrated copper (II) sulfate in Investigation 5? Explain.

INVESTIGATION 6: ATOMIC THEORY AND MATTER

Name_____

Section_____

WRITTEN EXERCISES--ASSUMPTIONS

In the study of chemistry we are forced to make many assumptions. For example, we must assume that our senses provide us with reliable information about the systems with which we work. We must assume that our measuring devices--rulers, graduated cylinders, balances, etc.--have been built with enough precision to allow us to collect valid data from them. You have been asked to assume that H_2O is a valid formula for water. It has been assumed that the greenish solid obtained from the burning of magnesium was Mg_3N_2, not MgN or Mg_2N_3.

All of the assumptions just referred to have one thing in common: We have done work in chemistry and advanced our understanding of chemistry by treating these assumptions as valid facts. Unless the assumptions we make are valid, they are of no use to us. We could assume, for instance, that the moon is made of green cheese and use that assumption to devise a wonderful plan to reduce the world's food shortage. We have learned from space exploration, however, that that assumption is probably not valid at all.

It should be clear that the assumptions that are valid are the ones that give us consistent results when we use them in our work.

Review the work that you have done so far this year in chemistry. In the space below, and on the back of this sheet, write down the assumptions that you have accepted as valid in order to do this work. Your statements should reflect a thoughtful review of each laboratory activity, class discussion, and demonstration so that you may illustrate all concepts and skills covered to this point in this course. Be as specific as you can be when listing the assumptions. Complete this activity before reading further.

You can see from your list that in order to make much progress in learning chemistry, you have to make many assumptions and assign some degree of validity to them. Considering assumptions as valid, however, implies that there are logical reasons for the assumptions. Look back at your list of assumptions. It would be helpful for you to make a list of reasons and then label your assumptions with the reasons that apply to them.

List of Reasons for Making Assumptions

EXAMPLE:

I accepted the authority of a textbook to validate this assumption.

INVESTIGATION 6: ATOMIC THEORY AND MATTER

Some preliminary assumptions that have been made this year have been based on data gathered in the laboratory. The name that we have given these assumptions is hypothesis. These hypotheses were then tested by experiment with one of two basic results. The experiment either validated our hypothesis (assumption) or forced us to make othEr assumptions. As long as our validated assumptions provided us with dependable results, we treated them as facts. We have analyzed these facts and combined them so that we have been able to establish a mental picture (model) for the way our systems work. This method of work is largely responsible for the historical development of chemistry and for the progress that is continually made in all areas of science.

We are ready to begin a study that relates directly to the most important assumption that we will make in the study of chemistry. The assumption is that ALL MATTER IS MADE UP OF ATOMS. In fact, there are several instances in which we have already used this assumption! For example, this assumption is the basis for writing specific formulas for compounds and balanced equations for chemical reactions. We need to examine this assumption in terms of established atomic theory and build a mental model for the atom that will support the results of the work we have done so far.

In order to accomplish the goal of building a workable model for the atom, we will have to rely heavily on the work of other investigators. The work that you have already done, however, can help you in the construction of this mental model and in the understanding of the way it works.

FIRST IONIZATION ENERGIES
(Last Column Relative to Hydrogen)

Atomic Number	Element Symbol	First Ionization Energy (volts)	Relative First Ionization Energy
1	H	13.53	1.00
2	He	24.46	1.81
3	Li	5.36	0.40
4	Be	9.28	0.69
5	B	8.26	0.61
6	C	11.22	0.83
7	N	14.48	1.07
8	O	13.55	1.00
9	F	17.34	1.28
10	Ne	21.47	1.59
11	Na	5.12	0.38
12	Mg	7.61	0.56
13	Al	5.96	0.44
14	Si	8.12	0.60
15	P	10.90	0.81
16	S	10.30	0.76
17	Cl	12.95	0.96
18	Ar	15.68	1.16
19	K	4.32	0.32
20	Ca	6.09	0.45
31	Ga	5.97	0.44
32	Ge	8.09	0.60
33	As	10.50	0.78
34	Se	9.70	0.72
35	Br	11.80	0.87
36	Kr	13.93	1.03
37	Rb	4.16	0.31
38	Sr	5.67	0.42
49	In	5.76	0.43
50	Sn	7.30	0.54
51	Sb	8.50	0.63
52	Te	8.96	0.66
53	I	10.60	0.78
54	Xe	12.08	0.89

SECOND RELATIVE IONIZATION ENERGIES

Atomic Number	Symbol	Second Relative Ionization Energy (relative to first ionization energy for H)
1	H	----
2	He	4.00
3	Li	5.56
4	Be	1.34
5	B	1.85
6	C	1.79
7	N	2.18
8	O	2.58
9	F	2.57
10	Ne	3.02
11	Na	3.48
12	Mg	1.11
13	Al	1.39
14	Si	1.20
15	P	1.45
16	S	1.72
17	Cl	1.75
18	Ar	2.05
19	K	2.34
20	Ca	1.27

COMMON POSITIVE IONS
(CATIONS)

Name	Symbol (with charge)	Name	Symbol (with charge)

COMMON NEGATIVE IONS (ANIONS)

Name	Symbol (with charge)	Name	Symbol (with charge)

INVESTIGATION 6: ATOMIC THEORY AND MATTER

EXPANSION OF THE IDEA

READING a: ATOMIC PUDDING

To develop a reasonable concept of the development of atomic theories, it is necessary to go far back in history to the time of the ancient Greeks. Yes, I know you get sick of hearing that the ancient Greeks took all of the good ideas (if you're not sick of it now, you will get that way), but that's just one of life's little inequities.

Around 430 B.C., a Greek named Democritus became the first known sponsor of an atomic theory. In fact, the word atom is Greek, and its literal meaning is indivisible. Many of the contemporaries of Democritus were very disturbed by the idea that substances could be infinitely divisible, or that any amount of material could be cut in half regardless of how tiny it was, if you had a sharp enough knife. To solve this problem, Democritus invented the smallest amount of matter that was possible, the indivisible particle, which translates to English as atom.

Democritus' theory failed to gain any wide degree of acceptance, but it was able to explain many things that were not explainable previously.

Aristotle, on the other hand, argued the notion that all matter is made up of combinations of various quantities of the elements. Aristotle reasoned that there were four basic elements--fire, water, earth, and air. There are also four prime qualities (hot, cold, dry, and moist), and every element has two prime qualities. Fire, for example, is dry and hot, and water is cold and moist. When one of the prime qualities is displaced by another, you have changed elements. If the cold in water is displaced by heat, the water is changed to air.

Aristotle's ideas of the elements were treated as truth (as were many of his other ideas), which had a tendency to interfere with the advancement of knowledge. When the truth is known, there is little need for further investigations, and his principle governed for the next 1500 years.

INVESTIGATION 6: ATOMIC THEORY AND MATTER

By the end of the 1700's, chemistry was in much the same position that physics had been in during Galileo's most productive years. The beliefs of Aristotle were being disproved by empirical observations, but there was no overall model or theory to explain what was being observed. Chemistry was becoming a quantitative science (based on numerical data) rather than a qualitative one (based on observations of properties and events). Chemists were finding that, although mixtures could be formed in any proportions, a chemical reaction had a precise recipe and any material that was extra would be left over after the reaction. For example, it became known about this time that combining 35 parts of chlorine by weight with 23 parts sodium by weight would produce table salt.

This set of measurements proved to be tailor-made for those who wished to promote an atomic theory of matter. One such person was the Italian scientist Amadeo Avogadro, who was able to convince the people around him of the usefulness of an atomic model of matter, but who failed to convince the vast majority of chemists.

The theory of the atomic composition of matter was given its first widespread acceptability by the English chemist (and school teacher) John Dalton. Dalton was able to convince chemists that an atomic theory could explain several regularities that were not adequately dealt with by other models. Dalton adapted the law of constant proportions by pointing out that the weight relationships (found empirically, as in the making of table salt) could be interpreted to indicate the relative weights of the atoms of various elements. Using this idea, Dalton calculated that sodium atoms were 23/35 as heavy as chlorine atoms, and that table salt was formed by combining one atom of sodium with one atom of chlorine. Dalton called this combination of atoms a molecule, which is the smallest division possible for a compound.

The atom helped to explain a lot of things, but it also left many facts unexplained. Many elements seemed to have more than one weight. Hydrogen, for example, was considered the lightest element, and water contained eight times as much oxygen as hydrogen (by weight). This 8:1 weight relationship seemed to conflict with other relationships between oxygen and other atoms, and between hydrogen and other atoms. It became obvious that a system composed entirely of diatomic compounds could not be made to fit the facts; some molecules must contain three or more atoms. Oxygen is 16 times as heavy as hydrogen, but in water there are two hydrogen atoms for every oxygen atom, giving the 8:1 weight relationship found by Dalton (and others).

The situation, however, remained very chaotic for about half a century. In 1858 a former student of Avogadro, the Italian chemist Stanislao Canizzaro published a book containing a consistent system of relative atomic weights. The book also gave the atomic composition of the better known (more common) compounds. This system enabled the Russian Dmitri Mendeleev to formulate the first periodic table in 1868. Chemistry at last had a workable atomic theory.

But the investigations continued. There was still a great amount of curiosity about what the atom looked like. Among those investigating the atom were J. J. Thomson, an Englishman who was director of the Cavendish Laboratories at Cambridge. Having found the mass of several types of ions, Thomson was experimenting with cathode rays, or what are found between two positive and negative electrodes in a vacuum tube. He found that cathode rays will pass through a magnetic field with remarkably little of the deflection (change of direction) that would be expected from a normal ion. Comparing the deflection of cathode rays to other ions (Thomson assumed cathode rays to be ions, as were many of the other rays or beams he worked with), Thomson found that cathode rays move at much higher speed than other ions. Further calculations showed that the cathode rays were different from ions in one of two ways: They carry 2,000 times as much charge as a hydrogen ion, or are about 1/2,000 as massive (heavy) as hydrogen ions. Another experiment solved the riddle of whether the cathode ray was a highly charged ion or a very tiny object. Thin sheets of material placed in the path of a cathode ray do not affect its motion. Thomson reasoned that if the cathode ray contained particles of atomic size, they should bounce off from a wall of closely packed atoms. On the other hand, if the particles were smaller than atoms, they could squeeze through the "cracks" between the atoms. Since these particles pass through, they must be sub-atomic in size. In 1897 Thomson announced the discovery of a new type of matter which he called the electron. This put an end to the original atomic theory in the Greek sense of the indivisible atom. The task for chemists and physicists then became to build a model of the parts of an atom and how they fit together.

To explain what the atom looked like after his discovery of the electron, in 1898 Thomson proposed what has come to be known as a plum pudding model (or watermelon, currant bun, or marbles-in-your-jello models) of the atom. Thomson's model envisioned the atom as being a positive cloud (the pudding) with enough electrons (the plums) in it to keep the atom neutral overall. The other forms of the model are very similar. The watermelon model, for example,

pictures the atom as a melon, with the meat being positive and the seeds the electrons. You'll have to figure the others out for yourself.

One of Thomson's students was a man by the name of Ernest Rutherford. He moved from New Zealand to England in 1895. While a student at Canterbury College he earned a scholarship to work in Thomson's laboratory. Three years later he accepted a professorship at McGill University in Montreal, where he spent nine years before returning to England to teach at the University of Manchester. It is Rutherford's work at Manchester that we are most interested in.

Rutherford noticed during his work with radiation--for which he won the 1908 Nobel prize--that alpha particles (helium atoms with two electrons removed) were deflected somewhat when they passed through thin sheets of mica. The deflection was small and might have been ignored, but Rutherford realized that this effect could not be adequately explained by a marbles-in-your-jello type of atomic model. The positive cloud (jello) would not change the path of the alpha particles (because the positiveness is so diffuse), and for the electrons to deflect the alpha particles would be roughly equivalent to changing the path of a bowling ball by rolling marbles at it.

Rutherford had a young assistant named Marsden, who needed a first project, so Rutherford set Marsden to work observing deflections of alpha particles as they passed through an extremely thin layer of gold. To do this Marsden sat for hours in a dark room letting his eyes adapt so that he could see the tiny flashes made by the alpha particles as they struck his viewing screen. He then would sit with the gold layer between himself and the alpha source and place his viewer at a certain angle of deflection (from the path the alpha particles would follow if they did not deflect at all). To establish the number of particles that were deflected to the angle at which he was observing, he would count the flashes on his viewing screen.

In general, the results were predictable. A great majority of the alpha particles was deflected a small amount, a matter of a few degrees. There were, however, a few particles (less than 1%) that changed direction by much greater angles, and a few particles even rebounded from the screen completely, coming off backwards.

These results were unsettling to Rutherford, so he asked Marsden's immediate supervisor, Hans Geiger (the inventor of the Geiger counter) to take an active role in the research. By 1911, enough quantitative data had been

collected to rule out the plum pudding model of the atom, which Rutherford replaced with another model.

Rutherford pictured the atom as a very dense, very tiny nucleus at the center (and containing the positive charge) surrounded by much empty space which holds a swarm of electrons. This model explains the Marsden-Geiger data by arguing that very few of the alpha particles that were shot through the gold layer actually got very close to the nucleus at all (in terms of atomic intranuclear distances) and so were deflected very little. A smaller proportion of particles would pass close enough to the nucleus to be deflected a greater amount, and an even smaller number of particles would "strike" the nucleus head-on, and be reflected backwards. In 1913 enough data had been gathered on the alpha particle deflection to estimate the size of the nucleus; smaller than 10^{-12} cm, compared to an atomic size of about 10^{-8} cm, or one ten-thousandth as large. (Actually, the gold nucleus is about 10^{-12} cm across, so Rutherford and his colleagues were in error when they judged it to be smaller than that.)

Meanwhile, back in the Cavendish Laboratories of J. J. Thomson, a new postdoctoral student had arrived from Denmark. Niels Bohr didn't really fit in at Cambridge, and described the laboratory as being in "a state of molecular chaos". His ideas were not well received, and when he met Rutherford at a reunion of Thomson's former students, Bohr was convinced that Rutherford's atomic model had better validity than Thomson's. At this point, the various histories conflict, but whether or not Bohr was ejected or left by mutual agreement (from Thomson's Cavendish laboratory), he soon found himself working with Rutherford at Manchester in the spring of 1912.

Bohr had an intuitive type of mind, and work by Max Planck and Albert Einstein hinted to him that there was a way to define the movement of electrons in Rutherford's atomic model that is more precise than an electron swarm. He enjoyed the atmosphere of the Manchester laboratories, but did not have time to work extensively on the problem. His scholarship had run out, and he returned to Copenhagen at the end of the spring. His marriage that summer and a first teaching job that fall prevented him from spending time on his puzzle.

In January 1913, a classmate visiting from England (H. M. Hansen) casually mentioned a formula with which Bohr was not familiar. When Bohr looked it up, the last piece of the puzzle fell into place. If the electrons orbited the nucleus in layers, Bohr reasoned, he could explain many things. The emission of light by gases when electric

current is passed through them (as in neon signs) could be explained, as could many mathematical questions. The theory proposed by Bohr places electrons in energy levels (layers), and to move an electron from one energy level to the next higher level requires that energy be added to the electron. For an electron to move from the energy level it is in to the next lower energy level requires that it give off energy (in some cases you can see this energy as light). Further, the electrons would always fill the lowest energy levels first, so electrons would give off energy until they had moved into the lowest possible level.

Just after Bohr's theory was proposed, World War I broke out, and many of the young scientists that gave the big research laboratories their main working force and productive hum were drafted to serve in opposing armies. For example, immediately after Oswald Moseley proved that x-rays could be explained through the use of the Bohr atom, he was conscripted and later died in the war at the battle of the Dardanelles.

As does any good theory, the Bohr atom raised many questions, and the period after the war was the time that these questions began to be asked, and answers began to be proposed and argued over. The theorists of the day were primarily younger men, and many of the liveliest arguments happened in places other than laboratories. Arnold Sommerfeld's Institut (a German word with obvious meaning) was near a beer hall (in Munich, Germany), and the waiters in the hall were under strict instructions not to clean the tables after their arguments. In fact, several journal articles supposedly were retrieved from the scribblings on the marble-topped tables there.

Further developments led to questions about the workings of the nucleus, and nuclear reactions began to be studied. Eventually a model was developed by Fritz Houtermans that explained the reactions which allow the sun to release such huge amounts of energy. On the day Houtermans completed his model, he had a date. Fritz and his date (whose name is lost) went for a walk down the quiet streets of Gottingen, and the lady remarked how pretty the stars were. Houtermans looked at her, and his reply catches the spirit of that era in science: "Yes, and would you have guessed that you are arm-in-arm with the only man alive who knows why they shine?"

EXPANSION THE IDEA

READING b: DOES THE AMOUNT OF MATTER MATTER?

One interesting (and sometimes mind-boggling) aspect of different areas of study is the specialized vocabulary. This is as true in languages, mathematics, and social studies as it is in the natural sciences. You have already become familiar with many words that have specific meanings and applications in chemical systems, and you will be exposed to many more. It is necessary, therefore, to develop a good understanding of these technical words in order to apply them properly. A helpful approach to good understanding, in many cases, is to find out how a word was invented or how an existing word came to be applied to a given situation. Without becoming too deeply involved in the background for the word, let's apply this approach to the word <u>isotopes.</u>

First of all, by definition, isotopes are atoms of the same element which have different weights. "But why the word 'isotopes'?", you might ask. We find that the prefix "iso-" came originally from the Greek language and means "same". The rest of the word "topes" is derived from the Greek word "topos" which means "place". We have, therefore, the literal translation of "same place" for our word "isotopes". (Read on further before you throw up your hands and scream, "It's all Greek to me!")

Historically, as the atomic theory of matter developed, the atom was considered to be the smallest particle of matter (therefore indivisible) and all atoms of a given element were thought to be <u>identical.</u> These quite reasonable assumptions, among others, were incorporated into the establishment of the early periodic tables of the elements. In a periodic table system of element classification, each element would occupy a particlar place on the table and each place should have a single element corresponding to it. As experimentation continued, however, people noticed what seemed to be two (sometimes more) kinds of elements that had the same chemical properties. This caused a big problem because it implied that more than one element could occupy a particular place on the table. We can see now that the reason for the problem was the identification of elements by their weights instead of some other criterion. Since the periodic system worked in so many other ways, the ultimate decision was to allow these two or more <u>elements</u> to occupy the same place on the table, hence the name isotopes. (This decision was made around 1913, not so long ago.)

As time went on, the discovery and classification of subatomic particles (that is, electron, proton and neutron) helped to explain the questions that had been raised by the discovery of isotopes. The definition of an element has now been changed so that the number of protons (not the weight) defines what the element is. The result of this definition is that isotopes no longer need to be defined as different elements in the same place. On the contrary, isotopes are shown to be atoms of the same element with different weights. Since the atomic weight is concentrated in the nucleus and since the number of protons is constant for a particular element, the difference in weight can only be accounted for by a difference in the number of neutrons. For example, the two isotopes of boron (with weights of 10 and 11) each have five protons in their nuclei. That is why they are boron and not some other element. The difference is that boron-10 has 5 neutrons and boron-11 has 6 neutrons.

The term <u>isotopes</u> is commonly used in relation to radioactive substances. It is important to notice, however, that not all isotopes are radioactive. In fact, most elements--boron, for example--have two or more naturally occurring, stable (non-radioactive) isotopes. This last statement gives a clue as to the meaning of <u>radioactive.</u> There are certain combinations of protons and neutrons in nuclei that tend to cause a disturbance of the strong, attractive forces of the nucleus. As a result of this disturbance, energy and/or particles are expelled from these nuclei with great force. The general result of this expulsion (or decay) is a nucleus that stays together better (is more stable). Depending on what is shot out of the nucleus, the element many be different than it was before. For example, alpha particles--composed of two protons and two neutrons--are expelled from uranium nuclei when uranium decays. It is clear that the remaining nucleus will weigh four atomic mass units less than it did before, and its atomic number will have been reduced by two protons. Since this nucleus now only contains 90 protons, it 's no longer uranium; it is now thorium.

The existence of isotopes of the various elements gives rise to the observed fact that most atomic weights are not whole numbers. They are better perceived as <u>weighted averages</u> of the isotopes that occur naturally. A naturally occurring sample of chlorine, for example, will contain some chlorine atoms that weigh 35 a.m.u. and some atoms that weigh 37 a.m.u. This information combined with the periodic table weight of chlorine (approximately 35.5) will permit us to determine the percentages of each of these isotopes in a natural sample of chlorine. The following calculation illustrates how these percentages can be determined. The

calculation will be based on a sample of 100 chlorine atoms because we are looking for parts for each one hundred (percentage) of each isotope. Our answers, therefore, can be read directly as percentages.

INFORMATION

1. An <u>average</u> atom weighs 35.5 a.m.u.
2. One hundred <u>average</u> atoms will weigh 100 times the weight of one, or 3550 a.m.u.
3. A certain number of those 100 atoms (X) will weigh 35 a.m.u. each, for a total weight of 35X a.m.u.
4. The rest of the atoms (100-X) will weigh 37 a.m.u. each, for a total weight of 37 (100-X) a.m.u.

NOTE: Be sure you understand this information before going on.

SOLUTION

The total weight of the atoms is 3550 a.m.u., which is the same as the <u>sum</u> of the weights contributed by atoms of each isotope. In equation form:
35X a.m.u.+37(100-X) a.m.u.=3550 a.m.u.

Then,

$$35X+3700-37X=3550$$

$$-2X=-150$$

$$X=75$$

Referring back to the information above, we see that "X" stands for the number of atoms that weigh 35 a.m.u. If 75 of our 100-atom sample is composed of atoms weighing 35 a.m.u. each, then the number of atoms weighing 37 a.m.u. each must be 25 (or 100-75).

The chemical reactions we have done (and will do in this course) utilize chemicals containing the various naturally occurring isotopes of the elements involved. It is apparent that the different isotopes behave the same chemically, because we have not had to adjust our reaction conditions or procedures to take all reactions into account. Chemically speaking, a chlorine is a chlorine is a chlorine, for all practical purposes. When discussing isotopes related to chemical change, therefore, it seems that a difference in the amount of matter doesn't really matter, except to find a more precise weight of an element in a system. There are occasions in some areas of endeavor, however, where the difference in the amount of matter (in isotopes of an element) <u>does</u> really matter.

INVESTIGATION 6: ATOMIC THEORY AND MATTER

Name_____

Section_____

READING 6: QUESTIONS

1. Given that the naturally occurring, stable isotopes of hydrogen have weights of 1 and 2, and that the table weight of hydrogen is 1.008, calculate the percentage of each isotope. Show your work.

2. Rubidium (Rb) has two isotopes, one of which weighs 85 a.m.u. If this isotope composes 72% of all rubidium, what is the weight of the other isotope (round to the nearest whole number). Use a table weight of 85.5.

INVESTIGATION 6: ATOMIC THEORY AND MATTER

EXPANSION OF THE IDEA

READING c: HOW BIG IS AN ATOM?

In answering the question that is the title to this article, someone might say "extremely small" or "the smallest particle of matter that can be distinguished as an element." When pressed for more specific information, someone else might say "an electron would be the size of a tennis ball and about a mile away from a nucleus that was the size of a golf ball." Both answers tell you something, but neither give you much information that could be useful in discussing atomic size (unless you wanted to know how far away an electron would be in an atom with a basketball-sized nucleus). A more specific answer is, "There would be about twenty-five million calcium atoms lined up in a row one centimeter long." This type of answer is of much greater value to us. We can now, at least, talk about how much of the one centimeter is contributed by each of the calcium atoms.

Let's assume, for the sake of discussion, that there <u>are</u> twenty-five million calcium atoms in a one centimeter row (atoms per cm). If we want to know the <u>diameter</u> of each atom, we have to find the inverse of the number we already have; that is, we want cm per atom. Inverting and dividing shows that each calcium atom would contribute forty billionths of a centimeter! As with many fantastically smaller or larger numbers used in scientific work, we need a convenient way of expressing the numbers so that we don't have to write 0.00000004 cm or even 4.0×10^{-8} cm to communicate atomic sizes. The convenience is provided in this case by defining a unit called the angstrom (A), where 1 angstrom equals 1×10^{-8} cm. This means, therefore, that the diameter of our calcium atom is 4 (A). There is no harm done by defining the angstrom unit because it represents a measurement that is so small that we can't really conceive of it. We can relate directy to centimeters and grams, but angstroms and atomic mass units are completely outside the realm of direct, sensual observation. We can, however, use the information to good advantage.

If we make the assumption that our atom is basically spherical in shape, then half the atomic diameter (the radius) will give us as much information as the diameter.

Adding electrons to an atom or taking them away will change the size of the atom, as well as the overall charge. The size change can be discussed just as well in terms of the change in radius as it can be in terms of a change in the diameter. The main reason, however, for using radii instead

of diameters is that chemical literature consistently discusses atomic and ionic sizes in terms of radii.

You have noticed patterns and trends in ionization energies for some of the A-column elements. Now you have the opportunity to look at some data for atomic radii to see if any patterns exist and what the meaning of any such patterns would be. Set up coordinates on graph paper, similar to the ones you set for plotting ionization energy data, with atomic numbers in ascending order on the X-axis. (Notice that the elements are listed in alphabetical order, not in numerical order.)

TABLE OF ATOMIC RADII FOR SOME A-COLUMN ELEMENTS

ELEMENT	ATOMIC RADIUS (A)
Aluminum	1.43
Argon	1.54
Boron	0.88
Carbon	0.77
Chlorine	0.99
Fluorine	0.64
Magnesium	1.60
Neon	1.12
Nitrogen	0.70
Oxygen	0.66
Phosphorus	1.10
Silicon	1.17
Sodium	1.86
Sulfur	1.04

Name_____

Section_____

READING C: QUESTIONS

After graphing the data in this reading, answer the following questions on the back side of your graph paper.

1. What trends and/or patterns do you see in the graph of these data?

2. Offer an explanation for the relationship between atomic number and atomic radius going from left to right in a row of elements.

3. How would the atomic radii of lithium and beryllium compare to (a) each other and (b) the rest of the second period radii? Explain.

4. How would the size of the radii of the elements in the fourth period (K, Ca, Ga, etc.) compare to the size of the corresponding elements in the third period (Na, Mg, Al, etc.)? Explain the difference in these radius sizes.

INVESTIGATION 6: ATOMIC THEORY AND MATTER

Name_____

Section_____

COMPOUND SHEET 1

Binary Compounds with A-Column Cations (Part 1)

A. Hydrogen

1. hydrogen fluoride
2. hydrogen chloride
3. hydrogen bromide
4. hydrogen iodide
5. hydrogen oxide
6. hydrogen sulfide
7. hydrogen nitride

B. Lithium

1. lithium fluoride
2. lithium chloride
3. lithium bromide
4. lithium iodide
5. lithium oxide
6. lithium sulfide
7. lithium nitride

C. Sodium

1. sodium fluoride
2. sodium chloride
3. sodium bromide
4. sodium iodide
5. sodium oxide
6. sodium sulfide
7. sodium nitride

D. Potassium

1. potassium fluoride
2. potassium chloride
3. potassium bromide
4. potassium iodide
5. potassium oxide
6. potassium sulfide
7. potassium nitride

C. Cesium

1. cesium fluoride
2. cesium chloride
3. cesium bromide
4. cesium iodide
5. cesium oxide
6. cesium sulfide
7. cesium nitride

F. Aluminum

1. aluminum fluoride
2. aluminum chloride
3. aluminum bromide
4. aluminum iodide
5. aluminum oxide
6. aluminum sulfide
7. aluminum nitride

Name_____

Section_____

COMPOUND SHEET 2

Binary Compounds with A-Column Cations (Part 2)

A. Magnesium

1. magnesium fluoride
2. magnesium chloride
3. magnesium bromide
4. magnesium iodide
5. magnesium oxide
6. magnesium sulfide
7. magnesium nitride

B. Calcium

1. calcium fluoride
2. calcium chloride
3. calcium bromide
4. calcium iodide
5. calcium oxide
6. calcium sulfide
7. calcium nitride

C. Barium

1. barium fluoride
2. barium chloride
3. barium bromide
4. barium iodide
5. barium oxide
6. barium sulfide
7. barium nitride

D. Test Your Understanding

1. boron fluoride
2. sodium phosphide
3. calcium arsenide
4. hydrogen telluride
5. rubidium selenide
6. beryllium bromide
7. strontium oxide
8. gallium chloride
9. magnesium phosphide
10. potassium selenide

E. Write the Names

1. H_2S
2. LiBr
3. NaF
4. KI
5. CsCl
6. Al_2S_3
7. Mg_3N_2
8. MgO
9. $CaCl_2$
10. BaS

Name_____

Section_____

COMPOUND SHEET 3

Binary Compounds With B-Column Cations (elements with only one common charge)

A. Zinc

1. zinc fluoride
2. zinc chloride
3. zinc bromide
4. zinc iodide
5. zinc oxide
6. zinc sulfide
7. zinc nitride

B. Silver

1. silver fluoride
2. silver chloride
3. silver bromide
4. silver iodide
5. silver oxide
6. silver sulfide
7. silver nitride

C. Lead

1. lead fluoride
2. lead chloride
3. lead bromide
4. lead iodide
5. lead oxide
6. lead sulfide
7. lead nitride

D. Test Your Understanding

1. zinc phosphide
2. zinc telluride
3. silver arsenide
4. silver selenide
5. lead arsenide
6. lead telluride

E. Write the Names

1. ZnO
2. ZnS
3. $AgCl$
4. Ag_2O
5. Pb_3N_2
6. PbI_2
7. AgI
8. Ag_3N
9. $ZnCl_2$
10. PbO

Name_____

Section_____

COMPOUND SHEET 4

Binary Compounds With B-Column Cations (elements with more than one common charge--Part 1)

A. Chromium (II)

1. chromium (II) fluoride
2. chromium (II) chloride
3. chromium (II) oxide
4. chromium (II) sulfide
5. chromium (II) nitride

B. Chromium (III)

1. chromium (III) bromide
2. chromium (III) iodide
3. chromium (III) oxide
4. chromium (III) sulfide
5. chromium (III) nitride

C. Iron (II)

1. iron (II) fluoride
2. iron (II) chloride
3. iron (II) oxide
4. iron (II) sulfide
5. iron (II) nitride

D. Iron (III)

1. iron (III) bromide
2. iron (III) iodide
3. iron (III) oxide
4. iron (III) sulfide
5. iron (III) nitride

E. Cobalt (II)

1. cobalt (II) fluoride
2. cobalt (II) chloride
3. cobalt (II) oxide
4. cobalt (II) sulfide
5. cobalt (II) nitride

F. Cobalt (III)

1. cobalt (III) bromide
2. cobalt (III) iodide
3. cobalt (III) oxide
4. cobalt (III) sulfide
5. cobalt (III) nitride

G. Nickel (II)

1. nickel (II) fluoride
2. nickel (II) chloride
3. nickel (II) oxide
4. nickel (II) sulfide
5. nickel (II) nitride

H. Nickel (III)

1. nickel (III) bromide
2. nickel (III) iodide
3. nickel (III) oxide
4. nickel (III) sulfide
5. nickel (III) nitride

Name_____

Section_____

COMPOUND SHEET 5

Binary Compounds with B-Column Cations (elements with more than one common charge--Part 2)

A. Copper (I)

1. copper (I) fluoride
2. copper (I) chloride
3. copper (I) oxide
4. copper (I) sulfide
5. copper (I) nitride

B. Copper (II)

1. copper (II) bromide
2. copper (II) iodide
3. copper (II) oxide
4. copper (II) sulfide
5. copper (II) nitride

C. Tin (II)

1. tin (II) fluoride
2. tin (II) chloride
3. tin (II) oxide
4. tin (II) sulfide
5. tin (II) nitride

D. Tin (IV)

1. tin (IV) bromide
2. tin (IV) iodide
3. tin (IV) oxide
4. tin (IV) sulfide
5. tin (IV) nitride

E. Mercury (I)

1. mercury (I) fluoride
2. mercury (I) chloride
3. mercury (I) oxide
4. mercury (I) sulfide

F. Mercury (II)

1. mercury (II) bromide
2. mercury (II) iodide
3. mercury (II) oxide
4. mercury (II) sulfide
5. mercury (II) nitride

G. Write the Names

1. CrO
2. Cr_2O_3
3. $CoCl_2$
4. $CuCl_2$
5. Cu_2S
6. SnF_2
7. SnO_2
8. FeO
9. Fe_2S_3
10. $NiCl_2$
11. Hg_2Cl_2
12. $HgCl_2$

INVESTIGATION 6: ATOMIC THEORY AND MATTER

Name_____

Section_____

COMPOUND SHEET 6

Review of Binary Compounds

A. Write the Formulas	B. Write the Names
1. hydrogen chloride	1. Hg_2Cl_2
2. hydrogen telluride	2. CaS
3. lithium oxide	3. LiBr
4. sodium bromide	4. ZnO
5. potassium chloride	5. Na_2O
6. cesium iodide	6. BaI_2
7. aluminum oxide	7. NiF_2
8. magnesium nitride	8. SnO_2
9. calcium oxide	9. $CoCl_2$
10. barium chloride	10. AlN
11. zinc sulfide	11. Fe_2S_3
12. silver sulfide	12. Na_2S
13. lead iodide	13. Na[illegible]
14. lead oxide	14. AgI
15. chromium (III) oxide	15. PbS
16. iron (II) sulfide	16. HF
17. iron (III) oxide	17. CsF
18. cobalt (III) nitride	18. SnF_2
19. nickel (II) sulfide	19. HgI_2
20. nickel (III) bromide	20. CuO
21. mercury (II) sulfide	21. H_2S
22. mercury (I) iodide	22. MgO
23. copper (I) chloride	23. Al_2S_3
24. tin (IV) chloride	24. CrF_2
25. copper (II) nitride	25. Cs_2O

INVESTIGATION 6: ATOMIC THEORY AND MATTER

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COMPOUND SHEET 7

Compounds Containing Anions in the Chlorate, Sulfate, and Nitrate Series

A. Chlorate

1. potassium chlorate
2. magnesium chlorate
3. nickel (II) chlorate
4. ammonium chlorate
5. aluminum chlorate

B. Chlorite

1. lithium chlorite
2. chromium (II) chlorite
3. cobalt (III) chlorite
4. mercury (II) chlorite
5. zinc chlorite

C. Hypochlorite

1. sodium hypochlorite
2. barium hypochlorite
3. silver hypochlorite
4. copper (I) hypochlorite
5. cesium hypochlorite

D. Perchlorate

1. calcium perchlorate
2. nickel (III) perchlorate
3. lithium perchlorate
4. tin (IV) perchlorate

E. Sulfate

1. sodium sulfate
2. barium sulfate
3. hydrogen sulfate
4. iron (III) sulfate
5. ammonium sulfate

F. Sulfite

1. silver sulfite
2. magnesium sulfite
3. tin (II) sulfite
4. aluminum sulfite
5. iron (II) sulfite

G. Nitrate

1. sodium nitrate
2. silver nitrate
3. copper (II) nitrate
4. chromium (III) nitrate
5. lead nitrate

H. Nitrite

1. potassium nitrite
2. aluminum nitrite
3. calcium nitrite
4. cobalt (II) nitrite

Name_____

Section_____

COMPOUND SHEET 8

Compounds with Polyatomic Anions that Contain Hydrogen

A. Monohydrogen Phosphate

1. lithium monohydrogen phosphate
2. calcium monohydrogen phosphate
3. nickel (III) monohydrogen phosphate
4. mercury (II) monohydrogen phosphate
5. copper (I) monohydrogen phosphate

B. Dihydrogen Phosphate

1. iron (III) dihydrogen phosphate
2. potassium dihydrogen phosphate
3. lead dihydrogen phosphate
4. magnesium dihydrogen phosphate
5. lithium dihydrogen phosphate

C. Bicarbonate

1. sodium bicarbonate
2. copper (II) bicarbonate
3. silver bicarbonate
4. nickel (III) bicarbonate
5. calcium bicarbonate

D. Bisulfate

1. chromium (II) bisulfate
2. aluminum bisulfate
3. zinc bisulfate
4. mercury (I) bisulfate
5. potassium bisulfate

E. Bisulfite

1. nickel (II) bisulfite
2. potassium bisulfite
3. chromium (III) bisulfite
4. copper (I) bisulfite
5. barium bisulfite

F. Bisulfide

1. zinc bisulfide
2. cesium bisulfide
3. iron (III) bisulfide
4. cobalt (II) bisulfide
5. silver bisulfide

G. Hydroxide

1. barium hydroxide
2. sodium hydroxide
3. tin (IV) hydroxide
4. chromium (III) hydroxide
5. calcium hydroxide

H. Acetate

1. copper (II) acetate
2. hydrogen acetate
3. ammonium acetate
4. chromium (III) acetate
5. silver acetate

Name_____

Section_____

COMPOUND SHEET 9

Compounds Containing Phosphate, Carbonate, Thiosulfate, Thiocyanate, Cyanide, Chromate, Dichromate, Oxalate, and Permanganate

A. Phosphate

1. potassium phosphate
2. magnesium phosphate
3. aluminum phosphate
4. sodium phosphate

B. Carbonate

1. sodium carbonate
2. calcium carbonate
3. copper (II) carbonate
4. chromium (III) carbonate

C. Thiosulfate

1. lithium thiosulfate
2. silver thiosulfate
3. zinc thiosulfate
4. cobalt (III) thiosulfate

D. Thiocyanate

1. mercury (II) thiocyanate
2. copper (I) thiocyanate
3. tin (IV) thiocyanate
4. potassium thiocyanate

E. Cyanide

1. silver cyanide
2. aluminum cyanide
3. barium cyanide
4. copper (I) cyanide

F. Chromate

1. zinc chromate
2. iron (III) chromate
3. sodium chromate
4. lead chromate

G. Dichromate

1. magnesium dichromate
2. potassium dichromate
3. cobalt (III) dichromate
4. tin (II) dichromate

H. Oxalate

1. copper (I) oxalate
2. hydrogen oxalate
3. barium oxalate
4. nickel (III) oxalate

I. Permanganate

1. potassium permanganate
2. iron (II) permanganate
3. chromium (III) permanganate
4. copper (I) permanganate

INVESTIGATION 6: ATOMIC THEORY AND MATTER

Name_____

Section_____

COMPOUND SHEET 10

Review of Compounds Containing Polyatomic Anions

A. Write the Formulas	B. Write the Names
1. chromium (III) permanganate	1. $KClO_3$
2. aluminum cyanide	2. $LiClO_2$
3. sodium phosphate	3. $NaClO$
4. copper (II) carbonate	4. $Ca(ClO_4)_2$
5. potassium dichromate	5. $BaSO_4$
6. barium oxalate	6. $Al_2(SO_3)_3$
7. copper (I) thiocyanate	7. $Pb(NO_3)_2$
8. lead chromate	8. $Ca(NO_2)_2$
9. zinc thiosulfate	9. $HgHPO_4$
10. ammonium chlorate	10. KH_2PO_4
11. silver sulfite	11. $NaHCO_3$
12. cesium hypochlorite	12. $Zn(HSO_4)_2$
13. potassium nitrite	13. $Cr(HSO_3)_3$
14. iron (III) sulfate	14. $CsHS$
15. mercury (II) chlorite	15. $Ca(OH)_2$
16. sodium nitrate	16. $NH_4C_2H_3O_2$
17. lithium perchlorate	17. $Mg_3(PO_4)_2$
18. calcium monohydrogen phosphate	18. $CaCO_3$
19. barium bisulfite	19. $Ag_2S_2O_3$
20. silver bicarbonate	20. $KSCN$
21. lithium dihydrogen phosphate	21. $Ba(CN)_2$
22. chromium (III) acetate	22. $ZnCrO_4$
23. tin (IV) hydroxide	23. $SnCr_2O_7$
24. zinc bisulfide	24. $Cu_2C_2O_4$
25. aluminum bisulfate	25. $KMnO_4$

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

Name_____

Section_____

INTRODUCTION

This investigation is the first of several that will involve the use of substances whose formulas can be produced from combinations of the common ions that you have learned. In doing this experiment, you will need to be very careful to follow the procedure and perform the weighings correctly.

EXPLORATION

1. Weigh 3.31 g of lead nitrate--$Pb(NO_3)_2$--into a clean, dry weighing cup. Transfer this solid to your clean 50 ml beaker. Add 25 ml of distilled water to the beaker and stir with your stirring rod until the solid is completely dissolved. Rinse off the stirring rod into the beaker.

2. Weigh 3.32 g of potassium iodide--KI--into a clean, dry weighing cup. Transfer this solid to your clean 100 ml beaker. Add 25 ml of distilled water to the beaker and stir with your stirring rod until the solid is completely dissolved. Rinse off the stirring rod into the beaker.

3. Pour the $Pb(NO_3)_2$ solution into the KI solution. Rinse the 50 ml beaker into the 100 ml beaker using your wash bottle.

4. Stir the mixture in your 100 ml beaker. Describe the results. Specify any evidence for a chemical change.

5. Weigh your clean, dry 400 ml beaker (to two decimal places).

_____

_____

_____

Average _____

6. Weigh a piece of filter paper to two decimal places.

_____

_____

_____

Average _____

7. Set up your funnel on a ring stand. Place your 400 ml beaker under the funnel to catch the filtrate, fold the filter paper and place it in the funnel, and then filter the mixture from the 100 ml beaker. Rinse any remaining residue from your 100 ml beaker into the filter, using your wash bottle.

8. When all the filtrate has passed through the filter, rinse the residue with distilled water from your wash bottle. Allow this rinse water to collect in your 400 ml beaker also.

9. Remove the filter paper and residue from the funnel and carefully place them in a clean 250 ml beaker. NOTE: BE CAREFUL NOT TO LOSE ANY OF THE RESIDUE.

10. Put the 250 ml beaker with the filter paper and residue in the place designated by your teacher.

11. Label your 400 ml beaker containing the filtrate and put it in the place designated by your teacher.

12. When the filter paper and residue are dry, weigh them to two decimal places.

_____

_____

_____

Average _____

13. Compute the weight of the residue. Show your calculations.

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

14. When your 400 ml beaker has been returned to you, weigh the system to two decimal places.

_____

_____

_____

Average _____

15. Record observations of the beaker's content. Compute the weight of the substance contained in the beaker. Show your calculations.

16. Record your data on the first line of the table and fill in the class data on the remaining lines.

Weight of Residue	Weight of Filtrate	% Weight Change

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

THE IDEA

1. State the evidence that a chemical change has occurred in this interaction.

2. List the formulas for the ions present in the lead nitrate [$Pb(NO_3)_2$] and potassium iodide [KI] systems before you mixed them.

3. List the formulas for the ions present after you mix the lead nitrate with the potassium iodide.

4. What change has occurred with the ions to explain the chemical change you noticed when you mixed lead nitrate and potassium iodide?

5. Write a chemical equation to represent the reaction studied in this investigation.

6. Compare the weights of the reactants with the weights of the products. What does this comparison illustrate about chemical reactions?

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

Name_____

Section_____

EXPANSION OF THE IDEA

ION EXCHANGE REACTIONS--EXPERIMENT

1. At your laboratory station you will find a set of six substances in labeled dropper bottles. Name each of the substances and identify the ions present.

Table II

	FORMULA	NAME	IONS
1			
2			
3			
4			
5			
6			

2. In the following procedures be careful not to contaminate the substances in the dropper bottles. (Use only the dropper provided with each bottle. Don't use the dropper to stir mixtures and don't allow the tip of the dropper to touch the mixtures.) Into a small clean test tube put a few drops of Substance Number 1. Add a few drops of Substance Number 2 and mix by agitating the test tube. Observe the mixture for any evidence of chemical reaction.

3. Making sure the test tubes you use are thoroughly clean and rinsed with distilled water, combine each of the six substances with the other five until you have tried all of the possible pairs. Enter your data in the following chart by indicating any evidence of chemical reaction you observe. Write NR (for no reaction) for any combination(s) which show no evidence of reaction.

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

Table III

	1	2	3	4	5	6
1						
2						
3						
4						
5						
6						

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

QUESTIONS

1. For each combination of substances which results in a precipitate write a balanced chemical equation.

2. For the reactions in Question 1, identify the product(s) precipitated. Eliminate other possibilities by examining the reactions which didn't result in a precipitate. Briefly state the logic used to identify each precipitate.

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM ICDIDE

EXPANSION OF THE IDEA

READING: PREDICTING THE PROPER PRODUCTS

In Investigation 7 and in earlier investigations, you were asked to predict the products of chemical reactions and then to write equations representing those reactions. In most cases you did this by writing the formulas for the reactants on the left side of an equation; then you proposed possible products using evidence that new materials were formed; and finally you used logic to account for any remaining substances.

Several kinds of evidence can be used to indicate that a chemical reaction has taken place. All of these kinds of evidence indicate that a different substance exists after the reaction than before. A change in the physical properties of substances can also indicate that a chemical reaction has taken place although it is not always easy to distinguish between a physical and a chemical change.

Specific products of chemical reactions can also be identified by their chemical and/or physical properties. The following is a discussion of some examples of chemical reaction indicators.

1. FORMATION OF A GAS. In Investigation 2, calcium was mixed with water. Bubbles were observed forming in the water. The gas was tested and found to be hydrogen gas. The reaction was:

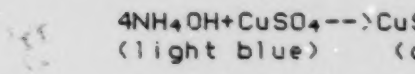
   $Ca+2H_2O \rightarrow Ca(OH)_2+H_2$.

   Following is a key to identifying some common gases.

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

NAME	FORMULA	COLOR	ODOR	COMBUSTION	MIXES WITH WATER	SPECIFIC TEST
Hydrogen	H_2	none	none	burns	no	burns explosively
Oxygen	O_2	none	none	supports burning	no	glowing splint flames
Carbon Dioxide	CO_2	none	none	no	yes	turns lime water milky
Nitrogen Dioxide	NO_2	brown	choking	no	yes	mixed with water turns litmus red
Bromine	Br_2	brown	choking	no	yes	liquid at room temperature
Chlorine	Cl_2	pale green	choking	no	yes	forms yellow mixture with water
Ammonia	NH_3	none	smelling salts	no	yes	mixed with water turns litmus blue
Sulfur Dioxide	SO_2	none	suffocating	no	yes	mixed with water turns litmus red
Hydrogen Sulfide	H_2S	none	rotten eggs	no	yes	---
Hydrogen Chloride	HCl	none	choking	no	yes	mixed with water turns litmus red
Water Vapor	H_2O	none	none	no	yes	cobalt chloride paper turns pink

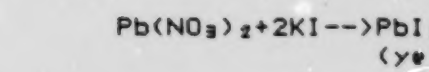
2. COLOR CHANGE. During Investigation 2, a reaction between ammonium hydroxide and copper (II) sulfate was carried out. Upon mixing, the substance turned from a pale blue to an intense blue. The color change indicates the formation of a new substance. The reaction is:



Many elements and ions are colored substances whose colors change when put in combination with different substances. Examples of such materials are the ions of elements from the middle (B-Column elements) of the periodic table (Cr, Mn, Fe, Co, Ni and Cu ions). Cu ions in compounds like $CuSO_4(H_2O)_5$, $Cu(NO_3)_2$, and $CuCl_2$ are light blue or green. Co ions are red in $CoCl_2(H_2O)_6$ and blue in $CoCl_2$. Mn ions are purple in $KMnO_4$ and black in MnO_2.

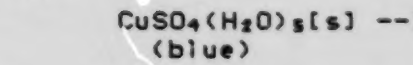
3. FORMATION OF A PRECIPITATE. In Investigation 7 lead nitrate and potassium iodide were mixed resulting in a substance which would not mix with water and thus precipitated. The precipitate was identified as lead iodide.

The overall reaction was:



As you gain experience with chemical systems, you will begin to recognize classes of compounds which mix with water and other classes which will not mix with water. Meanwhile, you will either have to rely on your memory or records to identify substances which will not mix with water.

4. A CHANGE IN WEIGHT. In Investigation 5, hydrated copper (II) sulfate was heated. The color of the solid faded from blue to white. Also, the weight of the residue was less than the weight of the reactant. The reaction was:



INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

In this reaction the change in weight is only one indication that a chemical reaction had taken place; there is also a color change and the evolution of a gas (H_2O vapor). However, sometimes a weight change is the only obvious indication of a chemical reaction. For example, heating white powdered hydrated barium chloride results in a white powder. However, the weight decreases. The reaction is:

$BaCl_2(H_2O)_2[s] \rightarrow BaCl_2[s] + 2H_2O[g]$

(white) 2.44 g → (white) 2.08 g

Of course, the total weight of the reactants should be equal to the weight of the products. However, since the balance of the weight (2.44 g-2.08g=0.36 g) is carried off as water vapor, it is not weighed.

By using the above indicators and other pieces of evidence (litmus test, conductivity, melting and boiling points and density measurements), you can make reasonable predictions concerning the products of a chemical reaction you might be studying in the laboratory. However, some products do not give any direct observation of their presence. They must be deduced indirectly. One of the most powerful methods of determining these products is to take advantage of the idea that <u>chemical reactions involve the rearrangement of atoms.</u> Because of this, there must be the same number and kind of atoms before a chemical reaction as there are after. By a process of elimination, you can often propose a complete set of products. For example, if you were to strongly heat potassium chlorate ($KClO_3$), a white residue results and a gas which flames a glowing splint is given off. The gas must be oxygen and you could write an equation:

$KClO_3[s] \rightarrow$ white residue $+ O_2[g]$

In support of this proposal, the weight of the residue is less than the weight of reactant. The equation as written doesn't account for the K and Cl atoms. You might be tempted to write the compound KClO as the residue since you will then account for all of the atoms with which you started.

$KClO_3[s] \rightarrow KClO[s]+O_2[g]$

You can test your hypothesis that the residue is KClO by measuring the melting point of the residue and looking up the melting point of KClO in a reference such as the Merck Index or the Handbook of Chemistry & Physics. The residue melts at 776 degrees Celsius. However, according to the Handbook of Chemistry & Physics, KClO decomposes when heated. Another possible compound which will account for K and Cl is KCl. The reaction would be:

$KClO_3[s] \rightarrow KCl[s]+O_2[g]$

KCl melts at 776 degrees Celsius and consequently is probably the residue. However, although the equation above accounts for the kinds of atoms, it does not account for the number. This requires balancing the equation by adjusting the coefficients.

$2KClO_3[s] \rightarrow 2KCl[s]+3O_2[g]$

The above reaction is called a decomposition reaction. Often recognizing types of reactions can aid in identifying possible products. A decomposition reaction starts with a single reactant which is divided into two or more products. A synthesis reaction does the opposite. Two or more reactants are combined to form a single product. In Investigation 5, the heating of magnesium is an example of a synthesis reaction.

$2Mg[s]+O_2[g] \rightarrow 2MgO[s]$

The reaction of Investigation 7 is:

$Pb(NO_3)_2+2KI \rightarrow PbI_2+2KNO_3$

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

The lead and potassium ions trade places to form the products. This kind of reaction is called a double displacement reaction. Another possible type of reaction involves a single displacement. Aluminum displaces hydrogen in this example from Investigation 2:

$2Al+6HCl \rightarrow 2AlCl_3+3H_2$

In summary, the following are steps to writing equations for reactions being studied in the laboratory.

1. Write correct formulas for the reactants and/or products you know exist.

2. Use chemical or physical evidence to infer other reactants or products. Write correct formulas for these.

3. Use logic to account for remaining atoms by proposing likely substances. Write correct formulas for these.

4. Balance the equation by adjusting the coefficients.

INVESTIGATION 7: LEAD NITRATE AND POTASSIUM IODIDE

Name_____

Section_____

READING 7

QUESTIONS

Predict the products and write balanced equations for each of the following reactions.

1. Magnesium metal reacts with hydrochloric acid (HCl) to produce a gas which burns explosively.

2. Lead acetate in water reacts with sodium iodide in water to produce a yellow precipitate. When the remaining liquid is boiled away, a white solid remains.

3. Silver nitrate in water reacts with potassium chromate in water to produce a red/brown precipitate.

4. When sulfur is heated in air, a suffocating, choking gas results.

5. Iron (II) sulfide reacts with sulfuric acid to produce a gas which smells like rotten eggs.

6. Lead metal reacts with sulfuric acid (H_2SO_4) to produce a white solid precipitate and a suffocating, choking gas.

7. Copper metal reacts with nitric acid (HNO_3) to produce a blue solution and a brown, choking gas.

INVESTIGATION 8: COMBINATIONS WITH ZINC

Name_____

Section_____

INTRODUCTION

In the last investigation, the accuracy of your weight data was stressed. This investigation uses less specific amounts of substances, but your careful observations are quite important.

EXPLORATION

1. Label a clean 18 x 150 mm test tube as System #1 and dispense approximately 10 ml of copper (II) sulfate solution into it.

2. Add one piece of mossy zinc metal to the $CuSO_4$ solution and observe the system carefully for 3 to 5 minutes. Record your observations.

3. Is there any evidence for a chemical change in System #1? If so, specify.

4. Slowly add 15 drops of concentrated sulfuric acid (H_2SO_4) to System #1, shake gently to mix, and record your careful observations.

5. Is there any evidence for a chemical change as a result of adding the acid to System #1? If so, specify.

INVESTIGATION 8: COMBINATIONS WITH ZINC

6. Set System #1 in your test tube rack and label a second clean 18 x 150 mm test tube as System #2. Dispense approximately 10 ml of copper (II) chloride solution into the System #2 test tube.

7. Add one piece of mossy zinc metal to the $CuCl_2$ solution and observe the system for 3 to 5 minutes. Record your observations.

8. Is there any evidence for a chemical change in System #2? If so, specify.

9. Slowly add 15 drops of concentrated hydrochloric acid (HCl) to System #2, shake gently to mix, and record your careful observations.

10. Is there any evidence for a chemical change as a result of adding the acid to System #2? If so, specify.

11. Set System #2 in your test tube rack. After the systems have been allowed to sit for about 10 minutes, record final observations on both systems.

System #1	System #2

THE IDEA

1. If you restrict yourself to the list of common ions, what products could possibly have been formed by the chemical changes you observed in System #1? Write equations for the reactions.

2. If you restrict yourself to the list of common ions, what products could possibly have been formed by the chemical changes you observed in System #2? Write equations for the reactions.

INVESTIGATION 8: COMBINATIONS WITH ZINC

EXPANSION OF THE IDEA

READING 8a--ACTIVE AND NOBLE METALS

If a study of chemical changes is to be possible, it stands to reason that the elements and compounds used in such a study must be changeable. Helium, for example, is so unreactive that it has never been known to combine chemically with itself or any other element. It is for this reason that helium is called an inert or noble gas. Nitroglycerin, on the other hand, is so chemically active that its behavior is unpredictable unless special precautions are taken when handling the substance. Furthermore, it is impossible (for practical purposes) to make direct observations of changes that occur during the reaction that nitroglycerin undergoes.

For our study of chemical changes to be most useful, we need to combine substances whose chemical activities lie between extremes such as those just mentioned.

The majority of the chemical elements are metals, and, as you might expect, some metals are more chemically active or noble than others. Consider the following ten metals: zinc, iron, nickel, tin, lead, copper, mercury, silver, gold and platinum. Every one of these metals will react chemically, so the term noble applied to any of them cannot mean the same thing as it does when applied to helium. The idea is clear, however, that the less active a metal is, the more noble (or chemically aloof) it is.

When trying to learn about a group of elements or compounds, it is often helpful to list several properties of those substances. Such a listing aids in observing trends, distinct similarities, or important differences among the substances. The following table gives information about each of the ten elements being considered. Study the table to see if there is any information given that can help you decide which of the metals are most active and which are least active.

INVESTIGATION 8: COMBINATIONS WITH ZINC

TABLE 1

Metal	M.P.(°C)	Density(g/ml)	Resistivity * (microohms-cm)
Zn	419.4	7.13	5.8
Fe	1540.	7.86	10.0
Ni	1453.	8.90	7.8
Sn	231.9	7.29	11.5
Pb	327.4	11.34	22.0
Cu	1083.	8.96	1.77
Hg	-38.9	13.55	95.8
Ag	961.	10.5	1.59
Au	1063.	19.3	2.44
Pt	1769.	21.45	10.0

*NOTE: This column tells us something about how well a metal conducts electricity. The smaller the number, the better the electrical conductivity.

It can be readily observed that there are no patterns developing among the columns of information; that is, the values do not generally increase, decrease or stay the same regardless of the order in which the elements are listed. Was it a waste of time to find out and record such information, then? The answer is an emphatic NO! This information may not give us the answer to the question we are asking, but it certainly has application to other questions that might be asked. For example, we see from the right-hand column that (for this grouping of metals) silver and copper are the best conductors of electricity, and mercury and lead are the poorest conductors. That information, combined with the information in the M.P. and Density columns for those metals, would help you decide which one would be best to use in a system that must be heavy, able to withstand high temperatures without melting, and able to conduct electricity well. (Which one would you choose?)

INVESTIGATION 8: COMBINATIONS WITH ZINC

Returning to the question of which metals are active and which are noble, we see that more information is needed. How are these metals found in nature? Isn't it reasonable that the more noble a metal is, the more likely it is to be found in its metallic state (called native metal)? Do these metals react differently under a given set of conditions? The following table should help you to answer these questions.

TABLE 2

Metal	Common Natural State	Reacts with cold, dilute HCl
Zn	sulfide ore	readily
Fe	oxide ore	readily
Ni	sulfide ore	slowly
Sn	oxide ore	very slowly
Pb	sulfide ore	almost none
Cu	native metal sulfide ore	none
Hg	sulfide ore native metal	none
Ag	native metal sulfide and other ores	none
Au	native metal	none
Pt	native metal arsenide ore	none

The trends that you can perceive by examining Table 2 give you a clear indication of the comparative chemical activities of these metals. These ten example metals are arranged in the same order that they would appear in a much larger listing of metals (and non-metals, for that matter) called the electrochemical series. In the larger electrochemical series there would be other metals falling between some of the metals appearing on our abbreviated list, but this shorter list allows us to compare the reactivities of the metals that are listed here.

INVESTIGATION 8: COMBINATIONS WITH ZINC

There is an interesting, experimentally observed phenomenon associated with some pairs of metals in the electrochemical series. For the purpose of describing this phenomenon, consider a water solution of mercury (II) chloride--$HgCl_2$. Clearly, this solution contains mercury that is chemically combined with another element (chlorine, in this example) so the mercury does not appear in its metallic state. If we add some solid, metallic iron to this solution, it won't be too long before we can observe tiny drops of liquid mercury collecting in the bottom of the container. Furthermore, if we weighed the iron metal before and after this experiment, we would find that the iron metal had <u>decreased</u> in weight. The conclusion to be drawn from these observations is that iron reacted chemically with the mercury in solution forcing the mercury out of solution and into its metallic (or elemental) state. In the process of this reaction, part of the iron--represented by the weight loss during the experiment--went into solution. Refer to Table 2 again and notice the relative positions of these two metals on the table.

Observations of the <u>common natural state</u> column of Table 2 indicates that most metals are found in compounds and not in their metallic state. In order to form these compounds (ores) the metals must have reacted chemically. We can consider, then, that isolating a metal in the metallic state involves <u>unreacting</u> the metal (or reversing the chemical reaction that allowed the compound to form). During the reaction between iron and mercury (II) chloride, the iron reacted to go into solution while the mercury reacted chemically to return to its metallic state. You have observed this many times in the corrosion of iron and copper. Since the iron (in the experiment described here) reacted and even forced mercury to react opposite to the natural reaction tendency, we say that iron is more reactive (or active) than mercury.

Numerous experiments that you have performed have illustrated the comparative chemical activities of copper and zinc. Compare their positions on Table 2. Reactions similar to the ones described here will also occur between mercury (II) chloride and zinc, tin and nickel. Some of these reactions occur more rapidly than others and, as you might predict, the closer the metal is to mercury, the slower the reaction will be (under a given set of conditions).

INVESTIGATION 8: COMBINATIONS WITH ZINC

Although these ideas of relative reactivity have many practical applications, they cannot be applied to economically isolating a metal on a grand scale. Other methods--such as electrolysis of melted compounds and decomposition of compounds by strong heating in an inert atmosphere--are more practical ways of obtaining large quantities of metals.

INVESTIGATION 8: COMBINATIONS WITH ZINC

Name_____

Section_____

READING 8a: QUESTIONS

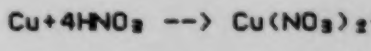
For purposes of these exercises; assume that hydrogen is always 1+ and oxygen is always 2-, <u>when they are part of a compound.</u>

1. Which elements are oxidized and which are reduced in the following reaction?

$$4HCl+MnO_2 \longrightarrow MnCl_2+2H_2O+CL_2$$

2. How many electrons are lost by the oxidations in the reaction above, and how many electrons are gained in the reductions?

3. List the spectator ions (name and formula) in the following reaction.



$$Cu+4HNO_3 \longrightarrow Cu(NO_3)_2+2NO_2+2H_2O$$

4. Is it possible for a spectator ion in an oxidation-reduction reaction to be oxidized or reduced? Explain your response.

5. Is it possible for the amount of oxidation to exceed the amount of reduction in a given oxidation-reduction reaction? Logically support your response.

INVESTIGATION 8: COMBINATIONS WITH ZINC

EXPANSION OF THE IDEA

READING 8b--ACTIVITY AND CHEMICAL REACTIONS

During this investigation, the relative ease with which metals gained or lost their electrons was related to the activity or reactivity of the metal. It is possible to react Zn metal with $Cu(NO_3)_2$. The reaction would be:

$$Zn+Cu(NO_3)_2 \longrightarrow Zn(NO_3)_2+Cu$$

However, the reverse reaction

$$Cu+Zn(NO_3)_2 \not\longrightarrow Cu(NO_3)_2+Zn$$

does not occur. This means that Zn loses electrons more readily than Cu and is thus more reactive or more easily oxidized. We can indicate Zn losing electrons by writing what is called a half-reaction:

$$Zn \longrightarrow 2e^-+Zn^{+2}$$

If we wanted to list half-reactions for zinc and copper in order of their ease of losing electrons, then our list would look like the following:

loses electrons most easily

$$Zn \longrightarrow 2e+Zn^{+2}$$

$$Cu \longrightarrow 2e+Cu^{+2}$$

This list of two half-reactions could be expanded using the ten metals discussed in Reading 8a by arranging the metals in a list in order of their ease of oxidation.

Materials other than metals can gain or lose electrons. For example:

$$2Cl^- \longrightarrow Cl_2+2e^-$$

Complex materials can also be oxidized. For example:

$$NO+2H_2O \longrightarrow NO_3^-+4H^++3e^-$$

INVESTIGATION 8: COMBINATIONS WITH ZINC

The ease with which these materials lose electrons can be compared with the metals considered before so that a more complete table could be constructed. This table appears below.

$Mg \rightarrow Mg^{+2}+2e^-$

$Al \rightarrow Al^{+3}+3e^-$

$Zn \rightarrow Zn^{+2}+2e^-$

$Fe \rightarrow Fe^{+2}+2e^-$

$Ni \rightarrow Ni^{+2}+2e^-$

$Sn \rightarrow Sn^{+2}+2e^-$

$Pb \rightarrow Pb^{+2} + 2e^-$

$H_2 \rightarrow 2H^++2e^-$

$Cu \rightarrow Cu^{+2}+2e^-$

$2I^- \rightarrow I_2+2e^-$

$Ag \rightarrow Ag^++1e^-$

$Hg \rightarrow Hg^{+2}+2e^-$

$NO+2H_2O \rightarrow NO_3^-+4H^++3e^-$

$2Br^- \rightarrow Br_2+2e^-$

$2H_2O \rightarrow O_2+4H^++4e^-$

$Mn^{+2}+2H_2O \rightarrow MnO_2+4H^++2e^-$

$2Br^- \rightarrow Br_2+2e^-$

$2H_2O \rightarrow O_2+4H^++4e^-$

$Mn^{+2}+2H_2O \rightarrow MnO_2+4H^++2e^-$

$2Cl^- \rightarrow Cl_2+2e^-$

$Au \rightarrow Au^{+3}+3e^-$

$2F^- \rightarrow F_2+2e^-$

INVESTIGATION 8: COMBINATIONS WITH ZINC

The above table can be used to predict chemical reactions. For example, does Cu react with I^-? From the table it can be seen that Cu loses electrons more easily than I^{-1} and that I_2 gains electrons more easily than Cu^{+2}. Therefore, the electrons lost by Cu can be gained by I_2 in the reaction

$$Cu + I_2 \longrightarrow Cu^{+2} + 2I^-$$
or
$$Cu + I_2 \longrightarrow CuI_2$$

As a rule of thumb, reactions can be developed by adding a half-reaction from the table with the reverse of a half-reaction below it. Therefore,

$$Cu \longrightarrow Cu^{+2} + [2e^-]$$
$$[2e^-] + I_2 \longrightarrow 2I^-$$

(cancel out the $[2e^-]$ at 1st and 2nd lines of upper equation)

$$Cu + I_2 \longrightarrow Cu^{+2} + 2I^-$$
or
$$Cu + I_2 \longrightarrow CuI_2$$

If the number of electrons does not cancel as above, then the coefficients of one or both half-reactions should be multiplied by a common multiple in order to cause the electrons to cancel. This will result in a balanced equation. As an example:

$$2(Al \longrightarrow Al^{+3} + 3e^-)$$
$$3(2e^- + I_2 \longrightarrow 2I^-)$$

$$2Al + 3I_2 \longrightarrow 2Al^{+3} + 6I^-$$
or
$$2Al + 3I_2 \longrightarrow 2AlI_3$$

INVESTIGATION 8: COMBINATIONS WITH ZINC

Equations formed by reversing the half-reaction nearest the top of the table will result in no reaction.

As an example, try this problem: Does Cu react with $AgNO_3$? Since $AgNO_3$ is made of Ag^+ and NO_3^-, the actual reaction is between $Cu+Ag^+$. From the half-reactions:

$$Cu \longrightarrow Cu^{+2}+2e^-$$

$$2(1e^-+Ag^+ \longrightarrow Ag)$$

$$Cu+2Ag^+ \longrightarrow Cu^{+2}+2Ag$$

Since the Ag half-reaction was reversed and is lower on the table, this reaction will occur. Since Ag^+ is found as $AgNO_3$, the reaction can be written:

$$Cu+2AgNO_3 \longrightarrow Ag\ Cu(NO_3)_2$$

The NO_3^- ion is a spectator ion in the reaction.

INVESTIGATION 8: COMBINATIONS WITH ZINC

Name_____

Section_____

READING 8b: QUESTIONS

Using the table in this reading, predict if the following will react.

1. $Ni+2HCl \rightarrow NiCl_2+H_2$.

2. $Cu+2HCl \rightarrow CuCl_2+H_2$.

3. $2Ag+I_2 \rightarrow 2AgI$.

Using the table in this reading, predict the products of the following reactions.

4. $Fe+SnBr_2 \rightarrow$

5. $Ag+Cl_2 \rightarrow$

6. $Sn+Fe(NO_3)_2 \rightarrow$

INVESTIGATION 8: COMBINATIONS WITH ZINC

Name_____

Section_____

PROBLEM SET 8

Balance the following equations using half-reactions.

1. $HNO_3 + H_2S \rightarrow NO + S + H_2O$.

2. $I_2 + HNO_3 \rightarrow HIO_3 + NO_2 + H_2O$.

3. $MnO_2 + HI \rightarrow MnI_2 + I_2 + H_2O$.

4. $NaI + H_2SO_4 \rightarrow H_2S + I_2 + Na_2SO_4 + H_2O$.

5. $NF_3 + AlCl_3 \rightarrow N_2 + Cl_2 + AlF_3$.

INVESTIGATION 8: COMBINATIONS WITH ZINC

6. $HNO_3 + H_3PO_3 \rightarrow NO + H_3PO_4 + H_2O$.

7. $I_2 + HClO + H_2O \rightarrow HIO_3 + HCl$.

8. $NH_3 + O_2 \rightarrow NO + H_2O$.

9. $H_2S + H_2SO_3 \rightarrow S + H_2O$.

10. $Cu_2S + HNO_3 \rightarrow Cu(NO_3)_2 + CuSO_4 + NO_2 + H_2O$.

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

Name_____

Section_____

INTRODUCTION

You have probably put hydrogen peroxide solution on a cut at one time or another. Do you recall the effect you observed? What caused the observed effect? Why is hydrogen peroxide normally stored in a brown bottle? Your reaction of hydrogen peroxide with manganese dioxide in this investigation will help in your understanding of the answers to these questions.

EXPLORATION

1. Clean and dry your 125 ml Erlenmeyer flask and determine its weight to two decimal places.

_____

_____

_____

average _____

2. Add 1 spoonful of manganese dioxide (MnO_2) to the flask and weigh the system.

_____

_____

_____

Average _____

3. Compute the weight of MnO_2 in the system. Show your calculations.

4. Add 15 ml of hydrogen peroxide (H_2O_2) solution to the system. Test the gas being produced with a glowing wood splint. Record your observations.

5. When the gas production ceases, label the flask and put it in the place designated by your teacher.

6. When the flask is returned to you, weigh the system again.

_____

_____

_____

average _____

7. Determine the weight of the contents of the flask. Show your calculations.

8. Record your observations of the contents of the flask.

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

THE IDEA

1. How does the weight of the contents of the flask compare with the weight of manganese dioxide that you put in the flask? (That is, what is the percentage weight change?)

2. From the information obtained in Step 7, what can you conclude about the identity of the contents of the flask? Explain.

3. How do you account for the production of oxygen gas in this reaction? Write an equation for the reaction.

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

Name_____

Section_____

EXPANSION OF THE IDEA

EXPERIMENT: H_2O_2 AND MnO_2 REVISITED

1. Set up the laboratory apparatus according to your teacher's instructions. Be sure that the rubber tubing is inserted in the eudiometer.

2. Weigh out the designated amount of MnO_2 on a piece of waxed paper. Record the weight. _____ g.

3. Transfer the MnO_2 to the reaction test tube. When you are ready to begin taking data, add 5 ml of H_2O_2 solution to the test tube, stopper it immediately, and record gas volumes every minute for 5 minutes. Record your data in the following table.

MnO_2 (g)	1 min. (ml)	2 min. (ml)	3 min. (ml)	4 min. (ml)	5 min. (ml)

4. At the end of the 5 minutes, dismantle your system and return the cleaned equipment to the place designated by your teacher.

5. Graph the volume data according to your teacher's instructions.

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

QUESTIONS

1. What general patterns do you see in the volume graphs?

2. Is there a relationship between the amount of catalyst and the rate of reaction? If so, state and explain the relationship. If not, explain why there is no relationship.

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET 9a: CATALYST PROBLEM

In the table below are generalized equations for six chemical reactions. The times at the right tell how long it takes for each reaction to occur. Each different letter in an equation stands for an element or compound, but a given letter will always represent the same compound. Substance B in Equation 1, for example, is the same substance as Substance B in Equations 2 through 6. The weights of the various substances in the reactions are written under the symbols for the substances.

Reactants		Products	Time (seconds)
1. A + B 5g 3g	——→	D + F + A 4g 3g 1g	58
2. A + B + H 5g 3g 1g	——→	D + F + A + H 4g 3g 1g 1g	45
3. A + B + B 1g 2g 1g	——→	D + J 2g 2g	90
4. A + B + B 1g 4g 1g	——→	D + J + B 2g 2g 2g	75
5. A + B + H 5g 3g 2g	——→	D + F + A + H 4g 3g 1g 2g	30
6. A + B 4g 3g	——→	D + F 4g 3g	45

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

1. Which reactions have a catalyst present? Support your choices with logical arguments.

2. You will notice that no two reactions listed in the table occur at the same rate. Account for the rate of each reaction in terms of one or more of the other reactions listed.

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET 9b: BALANCING CHEMICAL EQUATIONS

Reaction Type

1. $Na + Cl_2 \rightarrow NaCl$
2. $Sb + H_2O \rightarrow Sb_2O_3 + H_2$
3. $Zn + CrCl_3 \rightarrow CrCl_2 + ZnCl_2$
4. $NH_3 \rightarrow N_2 + H_2$
5. $P_2O_5 \rightarrow P_4O_6 + O_2$
6. $Na_2O \rightarrow Na + O_2$
7. $Ac(OH)_3 \rightarrow Ac_2O_3 + H_2O$
8. $MgO + H_2SO_4 \rightarrow MgSO_4 + H_2O$
9. $KClO_3 \rightarrow KCl + O_2$
10. $P_2O_5 + H_2O \rightarrow H_3PO_4$
11. $BaCl_2 + Na_2SO_4 \rightarrow BaSO_4 + NaCl$
12. $(NH_4)_2Cr_2O_7 \rightarrow N_2 + Cr_2O_3 + H_2O$
13. $IBr + NH_3 \rightarrow NI_3 + NH_4Br$
14. $Al + Cr_2O_3 \rightarrow Al_2O_3 + Cr$
15. $ZnS + O_2 \rightarrow ZnO + SO_2$
16. $F_2 + NaOH \rightarrow O_2 + NaF + H_2O$
17. $FeS_2 + O_2 \rightarrow Fe_2O_3 + SO_2$
18. $Cu_2S + HNO_3 \rightarrow Cu(NO_3)_2 + CuSO_4 + NO_2 + H_2O$
19. $MnCl_2 + Br_2 + NH_4OH \rightarrow MnO_2 + NH_4Cl + NH_4Br + H_2O$
20. $C_6H_{12} + O_2 \rightarrow CO_2 + H_2O$

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET 9c: BALANCING CHEMICAL EQUATIONS

Reaction Type

1. $Na + Cl_2 \rightarrow NaCl$
2. $FeO + O_2 \rightarrow Fe_2O_3$
3. $Rb + F_2 \rightarrow RbF$
4. $Hg(NO_3)_2 + NH_4SCN \rightarrow Hg(SCN)_2 + NH_4NO_3$
5. $Al + H_2SO_4 \rightarrow Al_2(SO_4)_3 + H_2$
6. $Zn + CrCl_3 \rightarrow CrCl_2 + ZnCl_2$
7. $FeO + O_2 \rightarrow Fe_2O_3$
8. $H_2 + Cl_2 \rightarrow HCl$
9. $FeCl_3 + NH_4OH \rightarrow Fe(OH)_3 + NH_4Cl$
10. $K + H_2O \rightarrow KOH + H_2$
11. $N_2 + H_2 \rightarrow NH_3$
12. $KNO_3 \rightarrow KNO_2 + O_2$
13. $NH_3 + O_2 \rightarrow N_2 + H_2O$
14. $KClO_3 \rightarrow O_2 + KCl$
15. $Cu_2S + O_2 \rightarrow Cu_2O + SO_2$
16. $P_2O_5 + H_2O \rightarrow H_3PO_4$
17. $FeS_2 + O_2 \rightarrow Fe_2O_3 + SO_2$
18. $C_2H_2 + O_2 \rightarrow CO_2 + H_2O$
19. $PbCl_2 + K_2CrO_4 \rightarrow KCl + PbCrO_4$
20. $Pb(C_2H_3O_2)_2 + K_2CrO_4 \rightarrow PbCrO_4 + KC_2H_3O_2$
21. $K_4Fe(CN)_6 + H_2SO_4 + H_2O \rightarrow K_2SO_4 + (NH_4)_2SO_4 + CO + FeSO_4$
22. $KrF_2 + H_2O \rightarrow Kr + O_2 + HF$
23. $C_6H_{12}O_6 \rightarrow C_2H_5OH + CO_2$
24. $CS_2 + Cl_2 \rightarrow CCl_4 + S_2Cl_2$
25. $C_3H_8 + CoF_3 \rightarrow C_3F_8 + HF + CoF_2$

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

EXPANSION OF THE IDEA

READING: ENZYMES--THE CATALYSTS OF LIVING SYSTEMS

The processes of life, such as digestion, respiration, and cell synthesis, involve thousands of chemical reactions. Each of the individual steps required to carry out these processes must take place smoothly, at the appropriate time, and at the temperature of the cell in which the reaction is necessary. In order for these steps to occur, the assistance of a catalyst is often necessary. A biological catalyst is called an enzyme. Thus, all complex organisms, including you, contain thousands of enzymes.

What is an enzyme? All of the enzymes which have been isolated have consisted chiefly of proteins, which are complex molecules made up of amino acids. Each kind of enzyme has a unique sequence of amino acids. The different kinds of enzymes are made possible by the large number of amino acids available and the numerous ways they can be arranged. If two enzymes catalyze different reactions, they will have different structures.

How do enzymes work? What is it about these very large, complex molecules that gives them the ability to catalyze reactions? These questions cannot be answered in atomic detail; but in general terms, chemists have learned how enzymes work. Like any catalyst, an enzyme has to do three things. One, it must be able to combine with the substance or substances (called substrates) on which it is to act. Secondly, the nature of the interaction with the substrate must be such that the reaction can proceed through a pathway that requires less energy to initiate the reaction. Thirdly, the enzyme must bind the reaction products relatively weakly. The products must be released quickly so that the next substrate can be bound and the reaction can proceed quickly.

A different enzyme is required for each substrate present. This ability of an enzyme to catalyze the reaction of only one kind of molecule is extremely important. This property of an enzyme helps preserve order in the functioning of living cell chemistry. To appreciate this, consider what would happen if one enzyme could catalyze reactions with all proteins. Chemical chaos would result, since this one enzyme would destroy the other enzymes (remember enzymes are made of protein) as well as the surrounding cells.

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

What is responsible for an enzyme's high degree of specificity? An enzyme molecule has a region called the active site which fits a specific substrate. The exact shape of the active site region of the enzyme must be designed to match, in a complimentary way, the shape of the substrate molecule. Enzyme molecules attach to the substrate molecules forming complexes. While the substrate and enzyme are in this complex, the bonds which hold the substrate together are weakened. This allows the reaction to occur with less energy and at the temperature of the cell. After the substrate bonds are broken, the enzyme separates from the substrate. Now the enzyme is free to combine with another subtrate molecule. This process leaves the enzyme unchanged and available for further combinations which explains why relatively small quantities of enzymes are necessary to catalyze reactions.

Enzymatic reactions are affected by several factors. The best temperature for enzyme activity is around 37°C which is the average body temperature. At temperatures of 50°C, the proteins which make up enzymes begin to break apart, thus changing the action of the enzyme. Another factor which affects enzymatic action is whether the environment is acidic, basic, or neutral. Most enzymes in your body function best in a neutral environment; however, the digestive enzyme, pepsin, is active in your stomach in acidic surroundings.

Inhibition of an enzymatic reaction is an important function to us. The action of enzymes can be inhibited when the active site of the enzyme combines with a substance other than the substrate. Many useful drugs exert their beneficial action by the inhibition of enzyme activity in bacteria. Bacteria such as staphylococcus require the simple compound para-aminobenzoic acid for their growth. They can grow and multiply in the human body because sufficient amounts of this compound occur in the blood and tissues. The compound sulfanilamide (a sulfa drug) was found to be effective in controlling many diseases caused by staphylococcus and other bacteria. The two compounds are very similar in size, shape, and chemical properties. The sulfanilamide fills the active site so the bacteria metabolic processes are interrupted. Usually fairly high concentrations of such a drug are needed for effective control of an infection, because the inhibitor should occupy as many active sites as possible. Also the natural substrate will have a greater affinity for the enzyme; thus, a great number of the false substrate inhibitor must be present. This is when the body's natural defenses can take over the bacteria and destroy them.

INVESTIGATION 9: HYDROGEN PEROXIDE AND MANGANESE DIOXIDE

Another example of enzyme inhibition is harmful to us. Mercuric chloride will inhibit the enzyme amylase in the breakdown of starch molecules. This takes place because the heavy metal ion reacts at an active site to form a stable compound. This explains why the metal ions such as mercury, lead, and silver are poisonous. If a small fraction of the enzymes in the body are inactivated, the effects may not be noticeable. However, if a substantial fraction of the enzyme is inactivated, metabolism will be upset and symptoms of poisoning occurs.

The characteristics of the biological catalysts, enzymes, that have been discussed also govern the action of catalysts in chemical reactions outside of living systems. Many industrial processes and chemical laboratory reactions depend on the catalytic process. The attempt to control automobile exhaust fumes with the use of a catalytic converter is an industrial application of catalysis. The exhaust gases, carbon monoxide and hydrocarbons, which are serious air pollutants, are channeled over the catalysts (metal oxides). The carbon monoxide and hydrocarbons are converted to harmless substances, carbon dioxide and water, then released into the air. Lead poisons the catalyst in the converter; therefore, lead-free gasoline must be used.

There are thousands of chemical reactions which are dependent on catalysis. Research is being conducted in an attempt to understand the enormous range of enzyme structures and the function of catalysts and enzymes.

INVESTIGATION 10: MASS RELATIONSHIPS

Name_____

Section_____

INTRODUCTION

This investigation involves a double displacement reaction in which you will determine one of the reactant weights, given information about the other reactant. It is important that you make <u>reasonable</u> predictions about the reactant weights. As an aid to this process, you may refer to the balanced equation for the reaction (Step 2) and to the experiment you did in Investigation 7. Care in handling the chemicals and care in weighing the systems will enhance your results.

EXPLORATION

1. Obtain about 1 ml of $AgNO_3$ solution and 1 ml NaCl solution in separate, clean test tubes. Combine the two solutions and record your observations.

2. Write a balanced equation for the chemical reaction observed.

3. Identify any precipitates formed in this reaction. Provide evidence to support your identification.

4. You will be using 20.0 ml of $AgNO_3$ solution labeled 100 g $AgNO_3$ per liter. How many grams of $AgNO_3$ will be in your 20.0 ml sample? Show your work.

5. Determine the number of grams of NaCl necessary to react with all of the $AgNO_3$ available in the reaction. Show your work below and check with your teacher before proceeding.

6. Determine the number of ml of NaCl solution corresponding to the weight of NaCl in Step 5. You will use a solution labeled 100 grams NaCl per liter. Show your calculations below.

7. Summarize the amounts of $AgNO_3$ and NaCl determined in the table below.

	weight	volume
$AgNO_3$		20.0 ml
NaCl		

8. Clean your 150 ml and 250 ml beakers with soap and water, rinse them with tap water, and then rinse the beakers with distilled water.

9. Dry and weigh the <u>250 ml beaker</u> to two decimal places.

_____

_____

_____

average _____

10. Clean two graduated cylinders with soap and water, rinse them with tap water, and then rinse the cylinders with distilled water.

    CAUTION: Silver nitrate solution will stain your skin and clothing. Thoroughly clean up <u>all</u> drips and spills.

11. Refer to Step 7 of this laboratory and obtain the amounts of $AgNO_3$ and NaCl solution you will use. Pour the two solutions together in the 150 ml beaker and stir with a clean stirring rod. Rinse each of the graduates into the beaker with a small amount of distilled water.

    NOTE: Be careful throughout the experiment to conserve all of the solid that is formed. If some is lost, be sure to make a note of the losses, so that these losses can be accounted for in the final results.

    NOTE: It is important to weigh your filter paper <u>before</u> filtering. Record the weight in Step 15.

12. Filter and rinse the solid AgCl. Collect all the filtrate in your 250 ml beaker. Carefully put the solid and the beaker containing the filtrate in the place designated by your teacher.

13. While the filtrate is drying, predict the amount of precipitate which will form and record your prediction on the table in Step 16. State the reasoning that led to your prediction.

14. Predict the amount of $NaNO_3$ which will be produced. State the reasoning that led to your prediction.

15. DATA

weight of AgCl and filter paper _____

weight of filter paper _____

weight of AgCl _____

16. Fill in the top line of the following table with the data you obtained. Obtain the data from the other groups in your class and complete the table.

WEIGHTS (grams)

$AgNO_3$	NaCl	(predicted) AgCl	(actual) AgCl	(predicted) $NaNO_3$	(actual) $NaNO_3$

INVESTIGATION 10: MASS RELATIONSHIPS

THE IDEA

1. Consider the various amounts of NaCl that were used by different groups. What conclusions can be drawn about the amount of NaCl and its effect on the amount of AgCl that was formed.

2. Compare the total weight of the reactants of this neaction with the total weight of the products. What conclusion can be drawn from this comparison?

3. In which of the systems were all of the reactants changed to products? Explain how you determined this.

INVESTIGATION 10: MASS RELATIONSHIPS

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET 10a: MOLAR CONVERSIONS

DIRECTIONS: Show all work. Use units. Answer to the nearest tenth.

1. How many grams are in 0.125 moles of Fe?

2. 90 grams represent how many moles of calcium?

3. How many grams are in 1.2 moles of gallium?

4. What is the weight of 2 moles of magnesium oxide?

5. 169 grams is equivalent to how many moles of cobalt (II) chloride?

6. How many grams are in 3.8 moles of magnesium phosphate?

7. How many grams are in 6.5 moles of sodium acetate?

INVESTIGATION 10: MASS RELATIONSHIPS

8. 324.7 grams of magnesium chlorate is how many moles?

9. How many grams are in 0.05 moles of aluminum hydroxide?

10. How much do 5.2 moles of strontium thiocyanate weigh?

INVESTIGATION 10: MASS RELATIONSHIPS

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET 10b: PERCENT COMPOSITION & SIMPLEST FORMULAS

1. Determine the percentage by weight of potassium and iodine in potassium iodide.

2. What is the percentage by weight of magnesium and oxygen in magnesium oxide?

3. Find the percentage by weight of each element in strontium chloride.

4. What is the percentage by weight of copper, carbon, and oxygen in copper(II)oxalate?

5. Determine the percentage by weight of the elements in ammonium hydroxide.

6. Find the percentage by weight of chromium, carbon, hydrogen, and oxygen in chromium(III)acetate.

7. A compound contains 36% nickel and 64% chlorine by weight. What is the simplest formula for this substance?

8. Find the simplest formula for boron carbide, a compound which is made of 78.6% boron and 21.4% carbon.

9. Calculate the formula for scandium nitrate which contains 19.5% Sc, 18.2% N, and 62.3% O.

10. A compound contains 78.6% boron and 21.4% hydrogen. What is its simplest formula?

INVESTIGATION 10: MASS RELATIONSHIPS

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET 10c: SIMPLEST FORMULAS

1. Acetic acid has the formula CH_3COOH. By weight what percentage of the compound is carbon? What percentage is oxygen? Hydrogen?

2. Krennerite contains 43.5% gold and the remainder is tellurium. What is its formula?

3. Hydroxylamine is 9% hydrogen, 48.5% oxygen and the remainder is nitrogen. What is its simplest formula?

4. A sample of bromodichloromethane contains 3 grams of carbon, 20 grams of bromine, 0.25 grams of hydrogen, and 17.7 grams of chlorine. What is its simplest formula?

5. Write the simplest formula for an amount of chrysoberyl composed of 2 grams of beryllium, 12 grams of aluminum, and 14 grams of oxygen ($Be_xAl_xO_x$).

INVESTIGATION 10: MASS RELATIONSHIPS

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET 10d: MOLE IN CHEMICAL REACTIONS

1. A mixture of hydrogen gas and fluorine gas would result in a violent reaction through which hydrogen fluoride is formed. (a) Write a balanced equation for the reaction. (b) How many moles of hydrogen fluoride will form from one mole of hydrogen gas with sufficient fluorine? (c) How many moles of fluorine are needed to produce 16 moles of hydrogen fluoride?

2. Crystals of potassium chlorate will decompose at high temperatures to produce oxygen gas and potassium chloride. (a) Write a balanced equation for this decomposition reaction. (b) How many moles of potassium chloride will be given by 0.66 moles of potassium chlorate? (c) How many grams of oxygen gas will be produced by 612.5 g of potassium chlorate?

3. The decomposition of calcium carbonate results in the formation of calcium oxide and carbon dioxide. (a) Write a balanced equation for this decomposition reaction. (b) How many moles of calcium oxide will be produced by 30 g of calcium carbonate? (c) How many moles of carbon dioxide will be formed by 212.0 g of calcium carbonate?

4. When magnesium reacts with dilute sulfuric acid, the products formed are magnesium sulfate and hydrogen. (a) Write a balanced equation for this reaction. (b) What kind of reaction is this? (c) How many grams of magnesium sulfate will be produced by 72 g of magnesium? (d) How many grams of hydrogen will be produced by 4.8 g of magnesium?

5. The electrolysis of molten aluminum oxide results in the production of aluminum and oxygen. (a) Write a balanced equation for the reaction. (b) How many moles of aluminum oxide will be required to produce 10.8 g of aluminum? (c) How many grams of oxygen will be produced if 10.8 g of aluminum are produced?

INVESTIGATION 10: MASS RELATIONSHIPS

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET 10e: REACTIONS WITH ONE REACTANT IN EXCESS

1. Ammonia gas (NH_3) can be manufactured by combining hydrogen and nitrogen gases. If 28 g of nitrogen and 7 g of hydrogen are available, how many grams of ammonia can be produced?

   Step 1. Write a balanced equation for the reaction.

   Step 2. Determine how many moles of each substance are available.

   Step 3. Which substance is in excess?

   Step 4. (a) How many moles of ammonia will be produced?
   (b) How many grams?

2. 16.6 g of sodium carbonate reacts with 9.49 g of HCl. Water and carbon dioxide are two of the three products formed. What will be the weight of the third product?

3. 92 g of magnesium are reacted with 511.25 g of copper(II)chloride. Elemental copper begins forming on the magnesium.

   (a) What type of chemical reaction is this? (Write a balanced equation.)

   (b) Which chemical will be completely used up at the end of the reaction?

   (c) How many grams of copper will be produced?

4. 17 g of silver nitrate are mixed with 17 g of NaCl. What will be the weight of the precipitate?

5. Given the reaction:

   $AlCl_3 + 4NaOH \longrightarrow NaAlO_2 + 3NaCl + 2H_2O$

   413.85 grams of $AlCl_3$ are reacted with an excess of NaOH. How many grams of $NaAlO_2$ will be produced by this reaction?

INVESTIGATION 10: MASS RELATIONSHIPS

EXPANSION OF THE IDEA

READING 10: YOU WON'T FIND ANY AVOGADRO IN A GUACA-MOLE SALAD

In 1802, J. L. Gay-Lussac discovered that all gases expand equally with an increase in temperature. In other words, one liter of chlorine gas will expand just as much as one liter of hydrogen gas if you increase the temperature of both equally.

The results of Gay-Lussac caught the attention of a young Italian lawyer named Amadeo Avogadro. In 1811, Avogadro published a paper containing his hypothesis to explain Gay-Lussac's equal expansion principle. Avogadro reasoned that if a liter of one gas expands as much as a liter of any other gas, then each liter of gas should have the same number of particles. Avogadro was careful to say that these particles didn't have to be atoms individually, but could be combinations of atoms. This became known as Avogadro's hypothesis and was widely ignored for half a century. Avogadro never knew of the wide acceptance his theory was to gain; he died in 1856.

In 1860, Stanislao Cannizzaro presented Avogadro's hypothesis at a large convention of chemists, and by then enough data had been gathered to make the hypothesis seem more reasonable to the rank-and-file chemists. It was too late for Avogadro to see it, but his hypothesis got the recognition it deserved, and is still called Avogadro's hypothesis and sometimes Avogadro's Law.

Since the general agreement after the convention of 1860 was that there were N molecules in one liter of oxygen gas and N molecules in one liter of hydrogen gas (all other things being equal), it gradually came into use that there was one mole (shortened from the German word molekulargewicht, molekular for molecular and gewicht for weight) of a substance in its equivalent weight (in grams). Oxygen, since it is diatomic, is found in the form O_2 and has an equivalent weight of 32 grams per mole of O_2 (16 grams per mole of O). About this time, experimental evidence also gave support to Avogadro. One equivalent weight of oxygen gas (one mole, 32 grams of O_2) takes up 22.4 liters (or 22,415 cubic centimeters, to be precise) of space under standard conditions and one mole of hydrogen gas (2 grams of H_2) also takes up 22.4 liters, again under standard conditions. Standard conditions are defined as 0 degrees Celsius and 1 atmosphere pressure.

INVESTIGATION 10: MASS RELATIONSHIPS

How many molecules are there in one mole of a substance? I can hear you asking that question even as I write. The first estimates of the number (to be known as Avogadro's Number) were based on the properties of gases at different temperatures and pressures. From these methods Loschmidt estimated Avogadro's Number to be 5 x 10^{23} which is admirably close to today's value.

In 1909, J. Perrin recalculated N (the symbol for Avogadro's Number) by using the motions of very small particles under the microscope. By taking many measurements and by using formulas I won't begin to force upon you, Perrin estimated the number of particles in one mole of a substance to be 7.2 x 10^{23}. After this, two more values for N followed very quickly.

Rutherford and Boltwood used radioactive decay observations to find their estimate of N in 1911. They obtained a value of 6.1 x 10^{23} particles per mole. In 1917, Millikan measured the charge on an electron, and this allowed him to refine Rutherford's estimate to a value of 6.07 x 10^{23}.

The most accurate value obtained to date (and the value accepted internationally) was obtained through measurements using the interactions of substances with X-rays and the densities of these substances. The International Union of Pure and Applied Chemistry (IUPAC) established 6.02252 x 10^{23} per mole as the official value for Avogadro's Number.

Some have objected that it is unfair to name the number of molecules in a mole after Avogadro, since he had no idea what the number was and had no part in its calculations. In defense of the Italian lawyer, however, he was considerably ahead of his time, and was the one who made the most important mental step in the process of creating a model to explain a previously unexplained observation and of creating a need for those calculations to be made.

As you can see, the mole was not an idea that was immediately accepted, and many estimations of its numerical value were necessary, even after the need for the concept was agreed upon. This process of advancing a hypothesis and using it over and over until enough data are gathered to build a better hypothesis is a pattern that is repeated countless times throughout the history of science.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

Name_____

Section_____

INTRODUCTION

There are many things about the dissolving process that can be observed, and you have already observed several of them in past investigations. In Investigation 1, for example, you studied the basic idea of dissolving and determined whether a substance was soluble in water or not. You also observed that the dissolving process affected conductivity in some cases as well as changing the effect on litmus paper. In this investigation you will study how the solution process affects temperature. What does temperature actually measure? What does the definition of temperature imply about what is happening in a dissolving process where the temperature changes?

EXPLORATION

1. Fill an 18 x 150 mm test tube half full with water and measure its temperature with a thermometer. Place two NaOH pellets in the water and agitate vigorously for about 15 seconds. Test the temperature of the solution and record your observations.

NOTE: Do not touch the NaOH pellets or the solution with your hands. If you spill the solution, wash immediately with running water and notify your teacher.

Repeat this procedure with fresh water using one spoonful of the following compounds. Record your observations.

2. NH_4NO_3

3. Anhydrous $MgSO_4$

4. Hydrated $MgSO_4$

5. Accurately weigh between 3 and 8 grams of $MgSO_4$ (anhydrous). Record the exact weight.

   Weight: _____ g.

6. Using a graduated cylinder add 100 ml of water to a polystyrene cup. Note the temperature of the water over several minutes until it is a constant value. Record the temperature.

   Temperature: _____ °C.

7. Add the $MgSO_4$ to the water while stirring vigorously with the thermometer. Keep a close check on the temperature of the solution. Record the highest temperature reached after all the $MgSO_4$ dissolves.

   Highest Temperature: _____ °C.

8. Find the change in temperature ($\triangle$ t). Show how you calculated this below.

9. Record your data on the table below. Obtain the data of the other students in your class and record the data on the table.

weight $MgSO_4$(g)	$\triangle$ t

10. Plot the weight of $MgSO_4$ versus change of temperature on a graph.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

THE IDEA

1. What conclusions can be drawn about dissolving processes from the data in Steps 1 through 4?

2. What trend is shown by the data on the graph?

3. Using your graph, compute the amount of heat that would be produced by the dissolving of 100 g of anhydrous $MgSO_4$.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

Name_____

Section_____

EXPANSION OF THE IDEA: TEMPERATURE AND CHEMICAL CHANGES

DEMONSTRATION

1. Describe what happens when solid $Ba(OH)_2(H_2O)_8$ is mixed with solid NH_4SCN.

2. Describe what happens when powdered Zn is mixed with $CuSO_4(H_2O)_5$ solution.

QUESTIONS

1. List evidence to indicate that the above interactions are chemical changes.

2. Write equations for the reactions.

3. Predict what the effect on the temperature would be if you varied the amounts of chemical reactants in this demonstration.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

Name_____

Section_____

EXPANSION OF THE IDEA: INPUT HEAT

DEMONSTRATION AND EXPERIMENT

1. Record your observations as your teacher lights a burner and heats $(NH_4)_2Cr_2O_7$.

2. Mix a half-spoonful of ammonium chloride and a half-spoonful of calcium hydroxide on a piece of paper. Transfer this mixture to an 18 x 150 mm test tube.

3. While holding the test tube with a wire test tube holder, warm the mixture gently with your burner and record your <u>careful</u> observations of the system.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

QUESTIONS

1. Discuss whether the reactions above are physical or chemical changes.

2. Discuss whether the reactions above are endothermic or exothermic changes.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

EXPANSION OF THE IDEA

READING 11a: MAKING THINGS HOT OR COLD

In Investigation 11, you have been exploring chemical and physical systems which gain or release energy in the form of heat. The heat was noticed as a change in the temperature of the reacting media. When NaOH was dissolved in water, for example, the heat it lost was absorbed by the water into which it was dissolved and the temperature rose. The dissolving of NaOH in water is said to be exothermic because it releases heat. If you were to have used two times as much NaOH, the amount of heat released would have doubled and the temperature change of the water would have also doubled.

Heat is often confused with temperature. To distinguish the two, consider the following experiment. If 5 g of NaOH were put into a beaker containing 50 ml of water, the temperature would be raised to 35°C. If a second beaker contained 100 ml of water, it would take 10 g of NaOH to raise the temperature to 35°C. Although the temperature is the same in both beakers, the container which has twice as much water has absorbed twice as much heat.

The amount of heat contained by an object is dependent on three factors. As shown in the example above, the weight of the object is one factor. Another factor is the temperature of the object. If you have two beakers which contain the same amount of water, the beaker which has the higher temperature contains the most heat.

The third and final factor which controls the amount of heat contained by an object is the material out of which the object is made. Different materials have a different ability to hold heat. This is called their heat capacity. Iron has less of a heat capacity than water. Using a constant heat source, it takes less time to heat 1 g of Fe to 85°C than to heat 1g of water to the same temperature. The iron requires less heat to increase its temperature than water does.

Heat is expressed in units called calories. One calorie is the amount of heat necessary to raise one gram of water one degree Celsius. The amount of heat necessary to raise 10 g of water from 25°C to 28°C is 30 calories.

(28°C-25°C=3°C; 3°Cx10 g water=30 calories).

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

If you dissolve 5.6 g of KOH in 100 g of water, the temperature of the water rises 14°C. The reason the temperature of the water rises is because the dissolving KOH gives off heat (that is, the dissolving of KOH is an exothermic process). If all of the heat given off by the KOH is absorbed by the water (an approximation which is usually a pretty good one in a system like this), then the heat given off by KOH should be equal to the heat gained by the water. Since the amount of heat gained by the water is 1400 calories (100 g waterx14°C), the amount of heat lost by the KOH must have also been 1400 calories. The assumption is that 5.6 g of KOH releases 1400 calories of heat.

This assumption is only good if all of the heat is transferred to the water. If some of it is transferred to the vessel containing the water or to the air above the vessel, then our value will have an error. If our vessel is made of a material which does not absorb heat readily and we do the reaction rapidly, then our assumption will be very close to being accurate. The vessel which contains the water is called a calorimeter, and although the polystyrene cups which you used during this investigation appear to be poor substitutes for <u>real</u> scientific equipment, they are actually some of the best calorimeters available.

Let us now return to the KOH dissolving. It was determined that 5.6 g of KOH released 1400 calories of heat when it dissolved. If more KOH was used, then more heat would be released. So the amount of heat would be 1400 calories per 5.6 g of KOH, or 1400 cal/5.6 g of KOH. The heat which is released is called the <u>heat of solution.</u> We can standardize this figure by dividing to find out how much heat is released per gram. This figure is 1400 cal/5.6 g=250 cal/g. The heat released by one mole of KOH is called the molar heat of solution for KOH. Since one mole of KOH weighs 56 g, the molar heat of solution for KOH is the amount of heat released when 56 g of KOH dissolves. The value is 250 cal/gx56 g/mole=14,000 cal or 14 kilocalories.

The amount of heat released or gained by a chemical reaction is called the heat of reaction. Since chemical equations express reactions in units of moles, the molar heat of reaction is a useful quantity. As an example, consider the following reaction which was carried out in a calorimeter containing 1000 g of water.

$C_2H_5OH+3O_2 \longrightarrow 2CO_2+3H_2O$

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

Five grams of C_2H_5OH were burned in excess oxygen. The temperature of the water started at 25°C and rose to 60°C.

Heat released = weight of water x temperature change ($\triangle$ t)

= 1000 g x35°C

= 35,000 cal of heat released

Since 5 g of C_2H_5OH has burned, the heat is:

$$\frac{35{,}000 \text{ cal}}{5\text{g } C_2H_5OH}$$

or

$$\frac{35{,}000 \text{ cal}}{5\text{g } C_2H_5OH} \times \frac{44\text{g}}{1 \text{ mole}} = 322{,}000 \text{ cal/mole}$$

The molar heat is 322 Kcal/mole.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

Name_____

Section_____

READING 11a: QUESTIONS

1. How much heat will be absorbed by 250 g of water when its temperature change is 35 degrees Celsius?

2. Dissolving a substance in 150 g of water released 75 calories. What was the temperature change of the water?

3. An endothermic reaction lowered the temperature of a weight of water by 13 degrees Celsius, while absorbing 39 kcal of heat. What is the weight of the water?

4. 100 g of water are placed into a calorimeter at a temperature of 21.3 degrees Celsius. 4.34 g of NH_4Cl are dissolved, during which the temperature of the water drops to 18.5 degrees Celsius. Compute the heat of solution for NH_4Cl in kcal/mole.

5. The molar heat of reaction for the following reaction is 673 kcal released per mole of $C_6H_{12}O_6$.

$$C_6H_{12}O_6[s]+6O_2[g]\longrightarrow 6CO_2[g]+6H_2O[l]$$

Calculate how much heat will be released if 10 g of $C_6H_{12}O_6$ are burned.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

EXPANSION OF THE IDEA

READING 11b: GREMLINS AND OILY RAGS

In most laboratory situations people attempt to make their systems as simple as possible. The main reason for doing this is to limit the number of variables in the system in order to control the experiment better. Data gathered from carefully controlled experiments can be used in developing mental conceptions (models) for the way systems work. The importance of models that are developed from experimental work is determined by how well those models can be applied to other systems, whether they are simple or complex systems. The more we are able to generalize with our models and the more we can apply our models to complex systems, the more important our models are.

You have experimented with a wide variety of chemical sytems, all of which have involved the transfer of energy during the reaction process. Some have given off energy in various forms (heat, light and sound), and others have absorbed energy in the reaction process. Common to most of the systems with which you have worked is a simplicity that has allowed you to write specific formulas for reactants, predict specific products, and write balanced equations for the interactions. In dealing with complex chemical systems we can look for observational evidence that will confirm (or deny) models that have been developed from the simpler systems. Observational evidence is quite important in working with the complex systems because it is difficult (often impossible) to write neat, clean, balanced equations for them.

Two such complex chemical interactions, which we will use in examining energy relationships in chemical change, are the frying of an egg and the burning of a piece of notebook paper. Notice that in both systems a chemical change will occur only after energy has been added to the systems. There are two important differences in these systems with respect to energy, however. First, it takes considerably less energy to start the egg frying reaction than is required to start the paper reaction. Second, when one stops adding energy to the egg, the reaction stops. By contrast, the paper reaction--once started--will continue after the initial energy source is removed. The reaction of the paper with air also gives off energy, primarily in the forms of heat and light.

Let's examine these two differences in more detail. One way to compare the amounts of energy necessary to start the reactions is to apply energy at the same rate to both systems. If, for example, we put the paper and an egg in the same pan and begin heating the pan, we observe that the egg soon begins changing chemically while the paper remains unchanged. Recall that the egg in Reading 2 changed chemically when the maximum temperature of the system was 100°C (the boiling point of water). This latter statement gives a somewhat more quantitative evaluation of the minimum energy required to make the egg react. We won't pursue, at this time, a more exact determination of the minimum energy necessary for reaction. The important point to be understood here is that there is a minimum amount of energy required in a system before a chemical reaction will occur. This minimum total energy of the reactants is called the activation energy. It is clear from our examples that the activation energy of the paper-air reaction is much higher than that of the egg reaction.

Without the addition of energy neither of the example reactions would occur. The second difference between the two systems, however, occurs after the activation energies for the reactions have been achieved. It has already been noted that the paper reaction continues and the egg reaction stops, upon removing the activating energy source. It is necessary that activation energy be available for a reaction to continue, so the source of that energy must be determined (after the initial source has been removed). This will only apply to the paper system because it is the one that continues reacting. We observe that a lot of energy is given off by the paper system as a result of the chemical change. It is reasonable to assume that some of the energy given off by the reaction could go into raising nearby reactants to the activation energy point. This reaction would continue until the system was cooled or until the reactants could no longer come in contact.

Interestingly, the amount of energy given off by the reaction is greater than the activation energy. If this were not true, the reaction of substances we call flammable --wood, oil, gas, etc.--would be of no value for supplying heat energy that can be used for various purposes. Since it is true, we must conclude that the products of the reaction contain less total energy than the reactants contained. This difference in energy can be accounted for in the energy given off by the reacting system. Such a reaction, during which heat is given off, is called an exothermic reaction.

What about the reaction in the egg system, then? We can say for sure that there is not enough energy given off during that reaction to continue to activate other reactant particles. (Recall that the reaction stops when the activating energy source is removed.) The most reasonable explanation for this observation is that energy is absorbed during this chemical change. In other words, the products have a greater total energy than the reactants in an endothermic reaction. If the egg reaction were exothermic like the paper reaction, we would have a miserable time trying to cook (or eat) a fried egg!

Spontaneous combustion is a strange (and often destructive) phenomenon which involves objects at low temperatures bursting into flame. When someone piles oily rags loosely in a poorly ventilated place, a good possibility exists for the occurrence of this phenomenon. This seems unreasonable, at first glance, because the activation energy for the reaction of oil and air or rag and air is much higher than that corresponding to room temperature. Let's examine the situation carefully and try to explain spontaneous combustion in terms of what we know about chemical changes.

The conditions under which spontaneous combustion occur are important. The rags are piled loosely, so there are many air spaces between the rags in the pile. This air is trapped in these spaces, which gives rise to the term dead air spaces. There is no air draft through the spaces, and since air is a poor conductor of heat, the lack of draft and the poor heat conduction property allow for heat generated in the air space to tend to remain there. It has been stated that room temperature does not provide the activation energy for the reaction. In general that is true, but review the idea of temperature again. Temperature measures the average kinetic energy of the particles in a system. Average values are typically composed of many values near the average, but a few values that are considerably higher and a few values that are considerably lower than the average. A few air and oil particles, therefore, will have enough energy to collide and achieve the activation energy for reaction. As you know, the reaction is exothermic, and if the heat does not escape, it will be absorbed by the system. The average kinetic energy of the region (air pocket) increases, logically resulting in a greater number of oil and air particles with enough energy to activate further reactions. If the heat from these reactions is produced faster than it is removed, we can easily see that it is just a matter of time before the reaction is occurring rapidly enough to produce a fire! The safety lesson to be learned from this discussion should be obvious.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

One other aspect of activation energy can be explored at this time. As you know, hydrogen peroxide decomposes so slowly at room temperature that it appears to be unreactive. When we added manganese dioxide, however, the reaction occurred quite rapidly with no net change in the manganese dioxide. Since the manganese dioxide was a catalyst and not a reactant, the chemical change that took place was the same as it would have been without the catalyst. The difference was that the reaction happened more rapidly at a lower energy level when the catalyst was present. We have seen reactions speed up before, but this was usually due to heating, not the addition of a catalyst. The <u>speeding up</u> factor is common to both situations, however. The conclusion could be drawn, therefore, that a catalyst has the effect of lowering the activation energy required for a reaction to occur.

INVESTIGATION 11: TEMPERATURE CHANGES AND DISSOLVING

Name_____

Section_____

READING 11b: QUESTIONS

1. Why is spun figerglass a better heat insulator than plate glass?

2. Compare the activation energies for the decomposition of hydrogen peroxide and the reaction of zinc with hydrochloric acid.

3. Why will the reaction between zinc and hydrochloric acid speed up when the system is heated?

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

Name_____

Section_____

INTRODUCTION

By the time you have finished with your work in this Investigation, you will probably be wondering how many ways water solutions of substances can possibly be studied. You may have observed that, in regard to natural systems, nothing is as simple as it seems to be at first glance. Water solutions of substances certainly follow that principle. The focus of the study in this investigation will be observations that you have made before. To focus your thinking while you are performing the laboratory work, look over the three questions in The Idea section of this Investigation.

EXPLORATION

1. In this Investigation you will be working with the eighteen compounds listed below. Fill in the blanks with correct formulas for the respective compounds.

(a)	Iron (II) sulfide	_____
(b)	Barium chloride	_____
(c)	Nickel (II) chloride	_____
(d)	Sodium fluoride	_____
(e)	Calcium chloride	_____
(f)	Copper (II) oxide	_____
(g)	Aluminum chloride	_____
(h)	Cobalt (II) chloride	_____
(i)	Potassium iodide	_____
(j)	Chromium (III) chloride	_____
(k)	Sodium bromide	_____
(l)	Magnesium chloride	_____
(m)	Lead oxide	_____
(n)	Calcium oxide	_____
(o)	Zinc sulfide	_____
(p)	Iron (III) oxide	_____
(q)	Strontium chloride	_____
(r)	Potassium bromide	_____

2. Test these compounds by completing the data table on the following page. To perform these tests, place a small amount of solid and 25 to 30 ml of distilled water in your 150 ml beaker.

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

Compound (formula)	Soluble in water	Solution conductive	Solution acidic, basic or neutral	Physical description of compound

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

THE IDEA

1. As a result of the tests you have performed, explain how you determined whether a compound was insoluble or soluble.

2. State any relationships that exist between physical appearance of the substance and their solubilities.

3. What changes do these compounds undergo as they are dissolved by water?

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

EXPANSION OF THE IDEA

READING 12: H_2O vs. THE SOLUTE

Water is often referred to as the universal solvent for good reasons. Almost every substance will dissolve in it, at least to a small degree. At the same time, many substances are more soluble in water than in any other common solvent.

Although crowned as the universal solvent, H_2O could not be named the absolute solvent. A number of substances elude its encompassing grasp. Others seem content with only a partial affinity for this chemical.

Are there principles which regulate water's attack upon a solute? Indeed there are, as there are also means by which we can predict whether a chemical will dissolve in water.

Imagine that you have placed a 30 gram crystal of ammonium chloride in 100 ml of water at 25°C (about room temperature). With some stirring, the mass of the solid decreases to 5 grams due to the action of the water on the solute. After many days of observation, the solid has changed shape, but its mass remains 5 grams. Have the molecules of ammonium chloride shifted in the crystal? This is unlikely since a solid is an object whose molecules assume relatively stationary positions (that is, they may wiggle somewhat, but they don't wander around).

Let us gather what data we have so far to produce a feasible explanation for this phenomenon.

(1) A maximum of 25 grams of NH_4Cl will dissolve under the existing conditions.

(2) The crystal's mass remains unchanged once it dwindles to 5 grams.

(3) The crystal changes shape.

Points one and two combine to describe what is known as a saturated solution. Saturated simply means that the solution contains the maximum amount of solute it can hold under a set of conditions. It would be logical to ask what characteristics of the dissolving process can be used to account for limits in the solubility of substances.

Ionic solids dissolve in water by breaking apart into ions, as you have already learned. It stands to reason that several water molecules would be required to overcome the attractive forces that hold the ions together in the solid state. Some ions, such as Ca^{2+} and O^{2-} would be more strongly attracted to each other than, say, Na^{+} and Cl^{-} would be. It follows, therefore, that more water molecules would be required to separate each CaO pair (and keep them separated) than would be required to dissolve each NaCl pair of ions. If this reasoning is valid, we could expect that more NaCl could dissolve than CaO in a given volume of water. Observations do, in fact, support those results! Different values for the solubility of various substances would, therefore, tell us something about the relative attractions between ions in compounds.

To deal with point number three, let us investigate what may be happening on the surface of the crystal. Water is a vigorous solvent. Its molecules are always in motion, and although it is saturated with ammonium and chloride ions, it continues to work diligently to drag off yet more ions. Occasionally water succeeds. But wait a minute. We noticed that our crystal remained 5 grams in mass even after many days. It is apparent that if weight is to be conserved, what goes up must come down. If some ions are going into solution, other ions must be coming out of solution, though not necessarily in the same place on the crystal.

Dynamic equilibrium is the technical name given for this give and take process. Although it is a term used to describe chemical systems, the principle involved is really familiar to all of us.

Football teams have a maximum of eleven players on the field at any one time. If a player is sent into the game from the bench, the rules of the game require that one of the eleven players must come off the field for the game to continue without penalty. In the same manner, when ammonium and chloride ions are dissolved, other ammonium and chloride ions must be redeposited on the crystal, given that the solution is saturated.

You may have noticed that careful note was made as to the exact temperature of the solution, that is, (25°C). From a practical standpoint, this factor is probably the most important variable in determining how much of the solute will dissolve before saturation occurs.

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

THE EFFECT OF TEMPERATURE ON SOLUBILITY

Within this subtitle is a word which has been defined already, though not named. Solubility refers to the amount of a substance which will dissolve in a given volume of water.

This quantity differs from substance to substance and is temperature dependent. Carefully examine the data in Table 1 below, paying attention to whether the compound is a solid or a gas. (The solubility of liquids will be dealt with at a later time.)

SOLUBILITY TABLE 1

Solubility is expressed in grams of solute per 100 ml of water/temperature of water in °C.

Formula	Cold Water	Hot Water
NH_3 (gas)	89.9 g/25°	7.4 g/100°
$Sr(NO_3)_2$ (solid)	8.41/0°	21.83/100°
CO_2 (gas)	0.348/0°	0.058/60°
$Na_2S_2O_3 \cdot 5H_2O$	79.4/0°	291.1/100°
PbI_2 (solid)	0.044/0°	0.41/100°
AgCl (solid)	0.000089/18°	0.0021/100°
NaCl (solid)	35.7/0°	39.12/100°
$(NH_4)_2Cr_2O_7$ (solid)	36.6/15°	89/30°
O_2 (gas)	0.007/0°	0.003/100°
$Ca(NO_3)_2$ (solid)	0.125/0°	0.077/100°
HCl (gas)	82.3/0°	56.1/60°
$AgC_2H_3O_2$ (solid)	1.02/20°	2.52/80°
CuO (solid)	insoluble	insoluble
Cl_2 (gas)	1.46/10°	0.561/30°

THE SOLUBILITY RANGE

After analyzing the chart, one can readily see that the solubility range is quite broad. At one extreme, CuO is considered to be completely insoluble. Water is totally unable to pry the copper away from the oxygen. Silver chloride is also classified as insoluble because only a small portion of it will go into solution. At another extreme, over 100 liters of ammonia gas can be dissolved in a liter of water at room temperature. Still a third category is recognized which falls between these two extremes.

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

Recall Investigation 2: Salt and Calcium. When calcium reacted with water, $Ca(OH)_2$ precipitated readily from the solution. This does not indicate that $Ca(OH)_2$ was totally insoluble. In fact, the solution above the $Ca(OH)_2$ solid *did* conduct electricity. If $Ca(OH)_2$ had been totally insoluble, no ions would have been available to conduct electricity. On the other hand, if calcium hydroxide were as soluble as ammonia, no precipitate would have formed. Substances that dissolve to a small degree, as shown by the behavior of $Ca(OH)_2$, are termed slightly soluble or partially soluble.

Sometimes the solubility range is more carefully subdivided by the following categories: insoluble, very slightly soluble, slightly soluble, soluble, freely soluble, very soluble. For our present purposes, the three divisions insoluble, slightly soluble, and soluble will be sufficiently descriptive.

PREDICTING SOLUBILITY

Patterns are much more easily remembered than particulars. Of course, we cannot expect every compound to fit every trend (did you notice which compound did not follow the observed trends in the first table?), but usually when a pattern is perceived, memorization is automatically diminished and retention increased. Examine carefully Solubility Table 2 for trends in solubility. Predicting solubility is made easier as patterns are more fully discerned.

SOLUBILITY TABLE 2

(S=soluble; Ss=slightly soluble; I=insoluble or nearly insoluble; D=decomposes in water; "—" means that the compound does not exist or is unstable.)

	ACETATE	BROMIDE	CARBONATE	CHLORIDE	CHROMATE	HYDROXIDE	IODIDE	NITRATE	OXIDE	PHOSPHATE	SULFATE	SULFIDE
Aluminum	S	S	—	S	S	I	S	S	I	I	S	D
Ammonium	S	S	S	S	S	S	S	S	—	S	S	S
Barium	S	S	I	S	S	Ss	S	S	Ss	I	I	D
Calcium	S	S	I	S	S	Ss	S	S	Ss	I	Ss	D
Copper (II)	S	S	I	S	S	I	—	S	I	I	S	I
Iron (II)	S	S	I	—	S	I	S	S	I	I	S	I
Iron (III)	—	S	—	—	S	I	—	S	I	I	S	D
Hydrogen	S	S	—	S	S	—	S	S	—	S	S	S
Lead	S	Ss	I	S	Ss	I	Ss	S	I	I	I	I
Magnesium	S	S	I	S	S	I	S	S	I	Ss	S	D
Mercury (I)	Ss	I	I	S	I	—	I	S	I	I	Ss	I
Mercury (II)	S	Ss	I	S	S	I	I	S	I	I	D	I
Potassium	S	S	S	S	S	S	S	S	S	S	S	S
Silver	Ss	I	I	S	I	—	I	S	I	I	Ss	I
Sodium	S	S	S	S	S	S	S	S	S	S	S	S
Strontium	S	S	I	S	S	Ss	S	S	Ss	I	Ss	D
Tin (II)	D	S	—	S	S	I	S	D	I	I	S	I
Zinc	S	S	I	S	S	I	S	S	I	I	S	I

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

Name_____

Section_____

READING 12: QUESTIONS

1. Referring to Table 1 of the Reading, state the general trend in solubility of <u>solids</u> as the temperature of the solvent is increased.

2. Referring to Table 1 of the Reading, state the general trend in solubility of <u>gases</u> as the temperature of the solvent is increased.

3. Suggest reasons for each of the above-stated trends. (Hint: Temperature is a measure of the average kinetic energy of the molecules. Higher temperatures, then, indicate higher kinetic energy.)

4. Make a study of Solubility Table 2 and use that information to fill in the following chart.

Ion	Soluble or Insoluble (Pattern)	Exceptions to the Pattern
Ammonium		
Acetate		
Bromide		
Chloride		
Iodide		
Oxide		
Sulfate		
Nitrate		
Carbonate		
Phosphate		
Potassium		
Sodium		

5. Given the following sets of reactants, predict the products, balance the equations, and use the patterns from Question 4 to predict whether a precipitate will form.

(a) $NH_4OH + Al(C_2H_3O_2)_3$ --->

precipitate_____

(b) $LiClO_4 + MgSO_4$ --->

precipitate_____

(c) $AgNO_3 + KI$ --->

precipitate_____

(d) $Pb(NO_3)_2 + H_3PO_4$ --->

precipitate_____

(e) $Ba(OH)_2 + CuSO_4$ --->

precipitate_____

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET 12a: SOLUTION CONCENTRATION EXERCISES

1. Calculate the molarity of 1.5 liters of a solution of $KClO_3$ containing 24.5 g of the solute.

2. What amount of ammonium dichromate would be needed to add to sufficient water to make one liter of 0.5 M solution?

3. Calculate the molarity of 0.17 g of barium hydroxide in 250 ml of solution.

4. How would you prepare 500 ml of a 0.7 M solution of sodium hydroxide?

5. How much KCl would you need to prepare 2.5 liters of 0.01 molar solution?

6. What is the molarity of acetic acid in vinegar which contains 5.0 grams of acetic acid per 100 ml of vinegar?

7. If 200 ml of a certain solution contains 28.2 grams of magnesium chloride, what is its molarity?

8. How many grams of solute would be contained in 5 liters of 0.25 M sodium chloride solution?

9. How many grams of solute would be contained in 450 ml of a 0.5 M solution of lithium perchlorate?

10. Formaldehyde (CH_2O) is a gas at room temperature. Its maximum solubility in water at room temperature and atmospheric pressure is 39.58 g per 100 ml of solution. What is the molar concentration of such a solution?

INVESTIGATION 12: PROPERTIES OF IONIC SUBSTANCES

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET 12b: SOLUTION CONCENTRATION EXERCISES

1. What is the molar concentration of 21.3 g of lithium perchlorate dissolved in enough water to make one liter of solution?

2. What volume of water would have to be added to the solution in Exercise 1 to make the concentration 0.001 M?

3. Given that the solubility of $Ca(OH)_2$ in water is 0.18 g/100 ml of solution, what would be the molar concentration of a saturated water solution of $Ca(OH)_2$?

4. How many grams of solute would be contained in 250 ml of a 6 M NaOH solution?

5. You know that zinc reacts with hydrochloric acid to produce hydrogen gas and zinc chloride. How many milliliters of a 4M HCl solution would be required to produce 1 mole of hydrogen gas, using this reaction?

6. The solubility of lead iodide is 0.05 g/100 ml of water. If 0.02 mole of lead nitrate is reacted with 0.04 mole of potassium iodide, what will be the minimum volume of solution required to keep a precipitate from forming?

7. A water soluble, liquid compound has a density of 0.8g/ml and a formula of CH_4O. What volume of this liquid would have to be added to water to make 1 liter of a 1.5 M solution?

8. The solubility of potassium iodide is 128 g/100 ml of water. Is it possible to make a water solution of KI with a concentration of 8M? Explain the reasoning that lead to your response.

9. Explain how you would make a 0.01 M solution of $CrCl_2$, using 61.5 g of the solid and all the water you needed.

10. Suppose that you are given a solution of silver nitrate containing 17 g $AgNO_3$ per liter of solution, and you are given a sodium chloride solution containing 11.7 g NaCl per liter of solution. How many liters of each solution would be needed to produce 0.6 moles of AgCl?

INVESTIGATION 13: ARRHENIUS SOLUTIONS

Name_____

Section_____

INTRODUCTION

You have noticed that certain of the compounds that you have tested in water solution have changed the color of litmus paper. What designation have we given to solutions that changed litmus paper from blue to pink?_____ What designation have we given to solutions that changed litmus paper from pink to blue?_____

In this Investigation you will have the opportunity to begin studying the source of the acid and base characteristics of solutions.

EXPLORATION

1. In the laboratory there are two solutions and two solids in labeled containers. Be sure that you are careful to avoid contaminating the solutions and solids when you are performing tests on them. Test each compound as a solid and as a solution with litmus paper and for conductivity.

2. Classify the compounds into acids and bases.

Acids	Bases

3. Look carefully at the formulas for the acids and the bases. Make a hypothesis of the characteristics that make a substance an acid or a base.

4. Using your hypothesis, predict whether the following compounds are acidic or basic. List only your predictions before proceeding to Step 5.

TABLE 1

Compounds	Predictions	Results of Conductivity Tests	Reaction to Litmus Paper
NH_4OH			
H_2SO_3			
HNO_3			
$Ba(OH)_2$			
KOH			
$HC_2H_3O_2$			
H_2SO_4			
$Ca(OH)_2$			
HBr			

5. In Table 1 you made predictions about which compounds were acids and which were bases. You may now test these predictions in the laboratory. List the tests you perform below and the results in Table 1.

INVESTIGATION 13: ARRHENIUS SOLUTIONS

THE IDEA

1. Is the presence of water necessary for a substance to exhibit acid or base characteristics? If so, what effect must water have on a substance to make it display those characteristics? If no, why not?

2. On the basis of your study, give <u>complete</u> definitions for <u>acids</u> and <u>bases</u>, including the criteria necessary to make the acid or base exhibit its characteristics.

INVESTIGATION 13: ARRHENIUS SOLUTIONS

Name_____

Section_____

EXPANSION OF THE IDEA: METALS IN ACIDS

EXPERIMENT

1. Write the chemical equation for the reaction of magnesium and hydrochloric acid.

2. What type of chemical reaction is this?

3. Set up your laboratory apparatus to collect a gas by displacement of water. You will read the volume of gas collected after each minute for a period of five minutes. First, add 10 ml of the acid to the test tube. When you are ready to collect the gas, add a 4 cm length of magnesium ribbon to the test tube. Each group will test three different acid solutions according to the teacher's instructions.

4. In Table 2, record the solutions your group will be testing and the data you collect by following the procedure from Step 3.

Table 2

10 ml of Various Acids	Volume of gas (in ml) after				
	1 minute	2 minutes	3 minutes	4 minutes	5 minutes

INVESTIGATION 12: ARRHENIUS SOLUTIONS

5. Record the class data in Table 3.

Table 3

Cycle Solution	Volume of gas (in ml) after				
	1 min	2 min	3 min	4 min	5 min
1.0 M HCl					
Average					
1.0 M $CaCO_3$					
Average					
1.2 M HCl					
Average					
1.2 M $CaCO_3$					
Average					
1.4 M HCl					
Average					
1.4 M $CaCO_3$					
Average					
1.6 M HCl					
Average					
1.6 M $CaCO_3$					
Average					

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INVESTIGATION 13: ARRHENIUS SOLUTIONS

QUESTIONS

1. Does the concentration of the acid solution affect the reaction rate? Explain.

2. Compare the reaction of magnesium with 0.4M HCl to the reaction of magnesium with 0.4M $HC_2H_3O_2$.
Offer an explanation for any differences in the rates of the reactions.

INVESTIGATION 13: ARRHENIUS SOLUTIONS

EXPANSION OF THE IDEA

READING 13a: IONS IN EQUILIBRIUM

This reading will concentrate on the behavior of two substances in water solution: ammonium chloride (NH_4Cl) and boric acid (H_3BO_3).

Both substances are soluble in water, even though ammonium chloride will dissolve to a greater extent than boric acid. (At 25°C, the solubility of ammonium chloride is 25 g/100 ml, and the solubility of boric acid is 5.2 g/100 ml.) In the dissolving process the compounds ionize. Ammonium chloride's ionization can be represented by:

$$NH_4Cl \longrightarrow NH_4^+ + Cl^-.$$

The ionization of boric acid has several possibilities, such as:

$$H_3BO_3 \longrightarrow H^+ + H_2BO_3^-$$

$$H_3BO_3 \longrightarrow 2H^+ + HBO_3^{2-}$$

$$H_3BO_3 \longrightarrow 3H^+ + BO_3^{3-}$$

A saturated solution (at 25°C) of ammonium chloride will conduct electricity (light bulb apparatus). This fact is not surprising given the model for the formation of charged particles in the solution upon dissolving. By contrast, a saturated solution (at 25°C) of boric acid will not conduct electricity (light bulb apparatus). An immediate thought might be that boric acid does not ionize. Certainly the conductivity test would support that idea, but recall that the litmus test shows that ionization does occur. Since the litmus changed from blue to pink, one or more of the ionization equations written above would be reasonable, indicating excess H^+ in the system.

Why is the boric acid solution non-conductive? Among the possible answers to this question is the argument that boric acid is less soluble than ammonium chloride; therefore, less ions will be produced in the boric acid solution. The weight of H_3BO_3 that will dissolve, under the conditions previously described, is approximately equal to 0.08 moles of boric acid.

Since we know that 25 g of NH_4Cl (0.47 moles) will dissolve under these conditions, we can be assured that 4.24 g (0.08 moles) of ammonium chloride will dissolve in 100 ml of water! We already know that such a boric acid solution will not conduct electricity. However, an ammonium chloride solution containing the same number of moles of solute (0.08 moles) will conduct.

What conclusion can be drawn from this test? The most apparent conclusion is that equal numbers of the two types of molecules (equal moles) do not produce equal numbers of ions when dissolved in water. Furthermore, the ammonium chloride solution contains the greater number of ions.

Assume that all of the ammonium chloride in our present example ionized. Recall that two ions--one ammonium ion and one chloride ion--are produced for each molecule of ammonium chloride that ionizes. When the 0.08 moles of ammonium chloride dissolve, therefore, we would obtain 0.16 moles of ions in the solution--enough to allow for light bulb conductivity. Applying this to the boric acid solution we must conclude that the first ionization equation for boric acid (above) is the only one that is reasonable, and even that does not happen for all of the H_3BO_3 that dissolves! (Do you see why?)

Now let's attempt to establish a logical explanation for these observations on boric acid. We can start by considering unionized boric acid molecules in the solution. Recall that water is a vigorous solvent, constantly tugging at the boric acid molecules. As a result of this pulling effect, some molecules ionize into H^+ and $H_2BO_3^-$ ions. We have seen that if this had happened to all the boric acid molecules, the solution would have conducted electricity. Since it doesn't conduct, this ionization process must stop somewhere short of total ionization. Reasonably, we could not expect the solution to reach a point where some sort of stop signal was given, after which there was no more activity. We would, instead, expect the vigorous pulling action of water to continue indefinitely because of the constant, random motion of the water molecules. This constant, random motion would continue to ionize boric acid molecules, but it would also permit H^+ and $H_2BO_3^-$ ions to move near each other, attract, and recombine into boric acid molecules. As the number of ions increases from the action of the water, the chances of ions recombining to form boric acid molecules also increase. We can conceive of a point in this interaction when the number of ions recombining at any time is equal to the number of molecules being pulled apart. At this point, which has been referred to before as dynamic

equilibrium, we see no net change in the number of ions or the number of boric acid molecules in the system.

Is it possible that the ammonium chloride solution does not totally ionize? Yes! The only thing we know for certain from our discussion here is that the ammonium chloride solution ionizes to a much greater extent than does the boric acid solution. We could reasonably assume that the ammonium chloride solution also establishes a dynamic equilibrium similar to that of the boric acid solution. This idea could be extended to all water solutions.

Whether a substance ionizes a little, a lot, or none at all is a matter whose importance depends on the system being considered. For example, the small amount of ionization in water can be ignored when light bulb conductivity is being considered. On the other hand, the ionization of water becomes an important consideration in the explanation of the acid-base characteristics of aluminium chloride or calcium oxide solutions.

EXPANSION OF THE IDEA

READING 13b: THOSE MIGHTY ACIDS

The stereotyped strong acid, in the minds of many individuals, is one which will burn through a desk top in a matter of seconds releasing great clouds of bilious smoke. After the acid has devoured the wood, it drips to the floor, only to destroy the floor beneath. Indeed, some acids are quite powerful in their action upon various substances. For instance, gold was considered to be a purely noble metal until aqua regia (royal water) came along. This solution of concentrated nitric and hydrochloric acids was so named because it was found to actually dissolve the royal metal, gold. Sulfuric acid, sometimes called the "King of Acids" (it is the most widely used industrial acid), is an excellent dehydrating agent. It's reaction with common table sugar, sucrose, is a fine example of how it readily draws water out of a substance. The once white crystal is transformed into a black carbon mass within a few seconds when subjected to a concentrated solution of this acid. (You might convince your teacher to demonstrate this reaction for you.)

Maybe you have wondered why some acids display such dramatic properties while others may seem to be virtually unreactive. Consider the effect on our bodies if common ascorbic acid (vitamin C) from citrus fruit were as reactive as aqua regia. Once ingested, imagine its attack upon sensitive throat linings. The fact is, however, that ascorbic acid is a weak acid, as are many other important

acids necessary for the maintenance of human life. Again, why is there such a marked difference in their chemical behaviors?

Historically, chemists have not answered this question with the same ease we might on paper. Much labor has gone into understanding not only the reactivity of acids (as well as bases), but also the composition of such substances. Chemical knowledge has not come overnight. The French chemist, Antoine Lavoisier (1743-1794), often called the father of modern chemisty, had stated that an acid had to contain oxygen. This assumption certainly proves true with nitric acid--HNO_3--and sulfuric acid--H_2SO_4, but not with hydrochloric acid--HCl. We may be tempted to say to Mr. Lavoisier, "Look at the formula--HCl; there is no oxygen in it." (This illustrates how much we can take for granted in our modern laboratories.)

The oxygen theory of acids prevailed until Humphry Davy (1778-1829), an Englishman noted for his discoveries of sodium, potassium, calcium, barium, magnesium, and strontium, questioned the Lavoisier hypothesis. Through investigation of muriatic acid (as HCl was then known), Davy determined that hydrogen and a gas he named chlorine were the components of this substance. Having refuted the previous hypothesis, he proposed a new hypothesis stating that acids were "particular compounds in which the hydrogen can be replaced by metals". This sounds more like our current definition, but we are not finished with our historical excursion yet.

A Swede, Svante Arrhenius, refined the model a bit further in 1884 while studying at the University of Uppsala. He proposed, after analyzing electrical conductivity in solutions, that the dissolved substance must break up into ions. Therefore, an acid now became regarded as a substance which released H^+ ions in solution. Similarly, a base was a substance which released OH^- ions. (Although Arrhenius' theory was not well accepted when first set forth, it earned him the Nobel prize in chemistry 19 years later.)

Why is the strength of one acid different from that of another? Sulfuric acid ionizes more readily than boric acid. It is not that boric acid is not soluble in water; it is! Something else must be the determining factor. Compare the sulfate ion with the borate ion. Which carries the greater charge? Which ion resists dissociation from hydrogen to a greater extent?

Sulfuric acid, for all practical purposes, ionizes

completely in its first dissociation,

$$H_2SO_4 \longrightarrow H^+ + HSO_4^-;$$

(that is, nearly every molecule undergoes this reaction.) Furthermore, about one out of ten of the remaining bisulfate ions dissociate again:

$$HSO_4^- \longrightarrow H^+ + SO_4^{2-}.$$

On the other hand, boric acid produces less than one dissociation in forty thousand molecules at room temperature in aqueous solution. This property contributes well to its use as an eye wash.

The degree of dissociation for an acid or a base is expressed through what is called the dissociation constant, designated K_A for an acid and K_B for a base.

(You may ask, why K? Remember C stands for Celsius. K seems to be next on the list of alphabetical possibilities.)

K_A for the first dissociation of carbonic acid is 4.30×10^{-7} at room temperature:

$$H_2CO_3 \longrightarrow H^+ + HCO_3^-.$$

If the molar concentration of the H^+ ion is multiplied by the molar concentration of the bicarbonate ion, and the product is divided by the molar concentration of molecular H_2CO_3 remaining in solution, one may obtan this constant.

$$\frac{[H^+] \times [HCO_3^-]}{[H_2CO_3]} = K_A$$

Tables 4 and 5 give values of K_A and K_B for several acids and bases.

TABLE 4

Equilibrium Constants
for some acids in
aqueous solutionS

COMPOUND	FORMULA	T°C	K_a
Acetic acid	$HC_2H_3O_2$	25	1.76×10^{-5}
Arsenic acid	H_3AsO_4	25	5.62×10^{-3}
Boric acid	H_3BO_3	25	5.79×10^{-10}
Carbonic acid	H_2CO_3	25	4.30×10^{-7}
Hydrochloric	HCl	25	very large
Hydrocyanic	HCN	25	4.93×10^{-10}
Nitric acid	HNO_3	25	very large
Nitrous acid	HNO_2	25	4.60×10^{-4}
Phosphoric acid	H_3PO_4	25	2.30×10^{-2}
Sulfuric acid	H_2SO_4	25	very large

TABLE 5

Equilibrium Constants for some bases in aqueous solutions

COMPOUND	FORMULA	T°C	K_b
Ammonium hydroxide	NH_4OH	25	1.79×10^{-5}
Calcium hydroxide	$Ca(OH)_2$	25	3.75×10^{-3}
Hydroxylamine	NH_2OH	20	1.07×10^{-8}
Lead hydroxide	$Pb(OH)_2$	25	9.60×10^{-4}
Potassium hydroxide	KOH	25	very large
Silver hydroxide	AgOH	25	1.10×10^{-4}
Sodium hydroxide	NaOH	25	very large
Zinc hydroxide	$Zn(OH)_2$	25	9.60×10^{-4}

INVESTIGATION 13: ARRHENIUS SOLUTIONS

Below are three examples of trends (experimentally determined) relating acid strength by family and period.

TREND 1: OXYACIDS WITHIN A FAMILY

$H_2SO_4 > H_2SeO_4 > H_2TeO_4$

As we analyze this series of oxyacids (acids containing oxygen), we notice that atomic size has increased from sulfur to tellurium. These elements in the respective acids take on a positive character when they are bonded in this manner with oxygen. Since repulsion of like charges increases over shorter distances, hydrogen is attracted to the sulfate group less than the selenate group. Certainly other factors are involved, but the phenomenon can be adequately explained by this effect of distance on dissociation. (Predict which is a stronger acid, perchloric acid--$HClO_4$--or periodic acid--HIO_4.)

TREND 2: OXYACIDS WITHIN A PERIOD

$HClO_4 > H_2SO_4 > H_3PO_4$

As in the previous series decreasing acidity can be linked partly with an increase in the size of the central atom. (Predict which acid is stronger, carbonic acid--H_2CO_3--or boric acid--H_3BO_3.)

TREND 3: HALOID ACIDS (an acid containing a halogen and no oxygen)

HI > HBr > HCl > HF

These acids are simple binary acids whose trend appears to contradict that observed for the VI-A elements which take on a positive character in an oxyacid when bonded to oxygen. Here the halogens (VII-A elements) possess a negative character. Which halogen has the largest atomic radius? Which of the haloid acids will have the longest bond length (that is, the greatest distance between atoms in the compound)? HI, of course!

Notice an interesting facet of the haloid family. Although hydriodic acid is the most readily ionized of this group and hydrofluoric acid the weakest, HF <u>cannot</u> be stored in glass bottles. The other three may be stored for years with no retaliation, but hydrofluoric acid must be stored in something like polyethylene plastic. Heinrich Schwanhard is credited for being the first to witness this rash effect of hydrogen fluoride on glass. In 1670, as he bent over an acid, his glasses seemed to become fogged with the vapor

that was evolving. When he attempted to wipe off the mist, he found that he could not because his glasses were actually etched by the substance. If, then, HF is more reactive with glass than HI, why is it considered a weaker acid? Let us return to structure again. Which holds on to hydrogen with greater attraction? HF. Which is the stronger acid, then? HI.

This seeming paradox serves well to bring us to the point that there are many facets to chemical reactivity. We cannot focus on one aspect of a chemical system without considering other important factors. In essence, we are pressed to further research, the heart of chemical study.

INVESTIGATION 13: ARRHENIUS SOLUTIONS

Name_____

Section_____

READING 13b: QUESTIONS

1. Order the acids and bases from Table 4 and 5 according to strength.

Acids	Bases

2. Write the dissociation equations and the expressions for K_A and K_B for an acid and a base.

3. Predict which is a stronger acid, perchloric acid ($HClO_4$) or periodic acid (HIO_4). Explain the reasoning used in your prediction.

4. Predict which acid is stronger, carbonic acid (H_2CO_3) or boric acid (H_3BO_3). Explain the reasoning used in your prediction.

INVESTIGATION 14: pH * ACIDS AND BASES

Name_____

Section_____

INTRODUCTION

A scale that is used to describe the extent to which a solution is acidic or basic is called the pH scale. Numerical values on the pH scale range from 0 to 14. Certain substances called indicators will show a given color inside the range of pH. Your task is to determine the range of the pH scale that corresponds to bases and the range of the pH scale that corresponds to acids, using indicators in base and acid solutions.

Nine indicators are listed below along with the colors they exhibit in ranges of pH.

Indicator	pH Range	Color
1. crystal violet	0 to 1.8	yellow
	above 1.8	blue
2. orange IV	0 to 2.8	red
	above 2.8	yellow
3. methyl orange	0 to 4.4	orange
	above 4.4	yellow
4. methyl red	0 to 6.0	red
	above 6.0	yellow
5. bromothymol blue	0 to 7.6	yellow
	above 7.6	blue
6. phenolphthalein	0 to 10	colorless
	above 10	red
7. alizarin yellow R	0 to 12	yellow
	above 12	red
8. azo violet	0 to 13	yellow
	above 13	violet

INVESTIGATION 14: pH * ACIDS AND BASES

EXPLORATION

1. Carefully dispense 10 ml of an acid or base solution into your clean 50 ml beaker. Your teacher will assign the solutions that your group is to test.

2. Add <u>two drops</u> of an indicator to the solution and swirl the beaker gently to mix. Compare the color obtained with the pH range listed for that indicator.

3. Pour this solution down the drain and rinse your beaker well with tap water, then rinse it with distilled water.

4. Repeat Steps 1 through 3 of this procedure using the same solution and different indicators. Determine the pH of your solution within a range of 2 pH units.

5. In the table below record the solution, indicator used, the indication given, and the information gained about the solution's pH. Record information for each test you perform.

SOLUTION	INDICATOR NAME	INDICATOR COLOR	pH INFORMATION

INVESTIGATION 14: pH * ACIDS AND BASES

THE IDEA

1. What pH range did the acid solution indicate?

2. What pH range did the base solutions indicate?

3. How does the idea of pH relate to the definition of acids and bases derived from Investigation 13?

INVESTIGATION 14: pH * ACIDS AND BASES

EXPANSION OF THE IDEA

READING: AN INTRODUCTION TO LeCHATELIER'S PRINCIPLE

What is meant by the expression ion product constant for water (Kw)? We say that Kw is the product of the hydrogen and hydroxide ion concentrations in pure water. You have hung additional information on the hook of this definition. You know, for example, that the H^+ and OH^- concentrations will be equal because it is only logical for water to break up (dissociate) in that way. After all, water is neither acidic or basic to litmus, and the presence of an excess of either of these ions would produce either an acidic or basic solution. You are also aware that this is a constant; that is, the value is dependable within a reasonable range of conditions.

A discussion of Kw takes on much greater meaning when we come to grips with the idea that water will dissociate into ions. But how does this happen? Although this dissociation is very slight, it happens spontaneously. There are not even enough ions produced to show conductivity with the light bulb conductivity apparatus. But why does water stop dissociating when the separate hydrogen and hydroxide ion concentrations reach a value of 1×10^{-7} moles/liter? To achieve reasonable answers to these questions, we need to re-examine our models of the behavior of molecules.

The kinetic molecular theory of matter states that all matter is in constant motion. You know from extensive observation of chemical reactions that reactants must come in contact with each other before a chemical reaction can occur. As you learned in READING 11b: GREMLINS AND OILY RAGS, reactants also must have attained a minimum amount of energy (activation energy) in order to react chemically. Now, let's see how we can combine these pieces of information to build a logical model for answering the questions that were asked earlier.

Imagine that a sample of water contains absolutely no ions. This is not the natural or usual situation for pure water, so some dissociation will have to occur in this imaginary sample. Since there is no reasonable way for molecules to simply fall apart, we can only assume that some energy will have to be added to the water molecules that will dissociate them into ions. In the absence of any outside disturbance (such as heat) the only reasonable way for a water molecule to gain energy would be for another water molecule to collide with it and transfer energy during the collision (much like the transfer of energy between colliding billiard balls). The increase in energy (all of

which need not be kinetic) can partially overcome the force of attraction between the atoms in the water molecule. The weakened bond can then be broken by the attraction of neighboring water molecules, resulting in the production of H^+ and OH^-. Using this simple model allows us to explain the presence of ions in pure water, but it doesn't give us a clue as to why the final ion concentrations are so low. At 24°C, the value for Kw is 1×10^{-14}, meaning that the concentration of each ion is 1×10^{-7} moles/liter. It follows that there are roughly 277 million molecules of water for every H^+ or OH^- ion in pure water. Certainly our model would not allow for complete ionization of water, but it would seem to allow for more than it does. One might even expect to get enough ionization to allow for light bulb conductivity.

Are there other reactions that behave in this way? (It is intriguing how attempting to answer one question so often leads to the asking of several other questions.) A great many reactions do not go to completion. To illustrate this idea of going to completion, recall the times when you have added salt to a glass of water (maybe because you had a sore throat, or you just needed a nice refreshing drink). If you had a tall glass full of water and you added one teaspoonful of salt, the dissolving of that salt went to completion (all of it dissolved). If, however, you had added much more (half a cup?), not all of the salt would dissolve regardless of how diligently you worked at dissolving it. The solution got to a point of saturation, stayed there, and the dissolving stopped; or did it?

You have learned, in READING 12: H_2O vs. THE SOLUTE, that experimental evidence shows how a solute continues to dissolve, even though the solution becomes saturated. The reason that we see no further change is that solute is leaving the solution (depositing as solid) just as rapidly as other solute particles are dissolving. We have a situation, then, in which the rate of dissolving equals the rate of deposition. This situation is called dynamic equilibrium; that is, the reactions occur at the same rate (equilibrium).

How can this deposition reaction occur? Since it is a reaction (or interaction) between sodium and chloride ions in the example above these ions must come together, and the fact of this coming together (collision) can be explained in terms of the constant, random motion of the particles in the solution. Now, the possibility exists for these ions colliding, sticking together, and depositing on the solid, beginning with the time that the first pair of ions dissolves. We have said that the rates are equal at

equilibrium for both reactions (dissolving and deposition). Initially, however, we are getting much more dissolving than deposition (the amount of solid salt decreases). This is reasonable, in terms of the collision model, because the smaller the number of ions in the solution, the less likely it is that two ions will collide, much less stay together and deposit. As the number of ions increases, the rate of collision and deposition increases until, at equilibrium, the rates for the two reactions are the same. Furthermore, these rates will remain the same until something is done to the system to cause the rates to get out of balance. When an imbalance occurs, the system will automatically move toward a new equilibrium position.

As an example of shifting to a new equilibrium point, think again about the saturated salt water solution with extra solid salt. Let's analyze what the effect would be of doubling the amount of water in this system. This cuts the concentration of ions in half immediately (same number of ions in twice the volume). The deposition rate decreases because the existing ions now occupy twice as much space, causing the rate of collisions to be greatly reduced. If we assume that the rate of dissolving remains the same (would anything have changed it?) that rate will now be greater than the new deposition rate and more of the solid will dissolve. You can readily see how this action will finally result in equilibrium being established again. The ion concentration increases, as more salt dissolves, and the increased ion concentrations allow the deposition reaction rate to increase, and so on to equilibrium.

A very interesting and important fact should be observed here. A reaction that does not go to completion will establish dynamic equilibrium at a particular point under a given set of reaction conditions. We often talk about the extent of a reaction with reference to how completely our original reactants combined to form products. When we observe the extent of the ionization reactions for pure water compared to the interaction of sodium chloride and water, it becomes immediately obvious that the sodium chloride ionization proceeds much farther to the right, according to the following equations.

$$H_2O \rightleftharpoons H^+ + OH^-$$

$$NaCl \rightleftharpoons Na^+ + Cl^-$$

Of course, it is assumed that the second reaction is performed in water. In fact, the parenthetical subscript (aq) could be added to each of the three components of the second equation to indicate the presence of water. Notice the double reaction arrows, indicating reactions in both directions.

Because of the consistency of reaction extent under given sets of conditions, the equilibrium concentrations of reactants and products can be determined and compared by ratio to arrive at an equilibrium constant. We will not go through a derivation except to say that the concentrations of products (right-hand side of the equation) are placed in the numerator and the concentrations of reactants (left-hand side of the equation) are placed in the denominator. If you consider this for a moment, it will become clear to you that the larger the value of the equilibrium constant, the farther the reaction goes toward completion before equilibrium is reached. It should be noted that the ionization constants for acids and bases (refer to INVESTIGATION 13, READING b: THOSE MIGHTY ACIDS) are, in fact, equilibrium constants. It would be helpful to review the formula and tables given in that article.

SUMMARY

Reactions that do not go to completion are said to establish a dynamic equlibrium. Since the extent of reacton will always be the same for a given set of reactants under constant conditions, an equilibrium constant can be determined for such reactions. The greater the value of this constant, the more toward completion the reaction goes. The essence of LeChatelier's Principle is that equilibrium can be upset by changing reaction conditions, but the system will always tend toward the re-establishment of chemical equilibrium.

INVESTIGATION 14: pH * ACIDS AND BASES

Name_____

Section_____

READING 14

QUESTIONS

1. According to LeChatelier's Principle, explain what would happen to the equilibrium in each of the following systems, when the stated changes are imposed on the systems.

   a. $Pb(NO_3)_2 + NaCl \rightarrow PbCl_2 + NaNO_3$, when more NaCl is added to the system.

   b. $N_2 + 3H_2 \rightarrow 2NH_3$ + heat (all are gases), when heat is added to the system.

   c. A saturated NaCl in water solution, when half the solution water is evaporated.

2. Referring to Reading b in Investigation 13: THOSE MIGHTY ACIDS, why would a larger value for the equilibrium constant indicate that a reaction would go farther toward completion than a reaction with a smaller value?

INVESTIGATION 15: ACID-BASE INTERACTION

Name_____

Section_____

INTRODUCTION

In Investigation 13 you studied the characteristics of substances that are called Arrhenius acids and bases. This investigation will deal with the effects of combining an Arrhenius acid with an Arrhenius base. The procedure is simple to perform, so you will be able to focus your full attention on developing an idea that will explain the interaction.

EXPLORATION

1. Put 10 ml of acetic acid solution in a clean 100 ml beaker. Add four drops of methyl orange indicator solution and swirl to mix.

2. What is the color of the resulting solution?_____

3. In what range of pH does the solution lie?_____

4. Add four drops of phenophthalein indicator solution to the mixture above and swirl to mix.

5. Does the addition of phenolphthalein have an effect on the color of this solution? Explain your response.

6. Put 7 ml of sodium hydroxide solution into a clean 50 ml beaker.

7. Using a clean dropper pipet <u>from your drawer,</u> slowly add three drops of sodium hydroxide solution to the acid solution and swirl to mix. Record your observations.

8. Continue adding sodium hydroxide solution <u>dropwise with swirling</u> until you have added all the sodium hydroxide solution. Write a detailed summary of your observations during this procedure. (NOTE: You may be wise to test the solution with litmus paper from time to time.)

INVESTIGATION 15: ACID-BASE INTERACTION

THE IDEA

1. Did this procedure involve a chemical change between the acid and base? Explain your response thoroughly. If so, write a balanced equation for the reaction.

2. What does the beaker contain after you have added three drops of NaOH solution to the acetic acid?

3. What does the beaker contain when you have added enough NaOH to cause the color to fade from orange to yellow?

4. What does the beaker contain when you have added enough NaOH to cause the color to change to a dark pink?

INVESTIGATION 15: ACID-BASE INTERACTION

Name_____

Section_____

EXPANSION OF THE IDEA: BARIUM HYDROXIDE + SULFURIC ACID

DEMONSTRATION

1. Diagram and label the experimental set-up used in this demonstration.

2. Describe what happens as barium hydroxide--$Ba(OH)_2$--is added to the sulfuric acid solution--H_2SO_4.

QUESTIONS

1. What evidence shows that this interaction is a chemical reaction?

2. Write a balanced chemical equation for the reaction.

3. At what point in this demonstration is there an equivalent amount of acid and base in the beaker? What is the pH of the solution at this point?

INVESTIGATION 15: ACID-BASE INTERACTION

EXPANSION OF THE IDEA

READING 15: DOES AN ACID-BASE NEUTRALIZATION RESULT IN A NEUTRAL SOLUTION?

Not necessarily. . . .

Now that the question has been answered, I suppose you want to know why. In order to develop the argument, it is necessary to review a few ideas you learned in earlier activities.

If you add an equivalent number of moles to a sample of base, the resulting mixture is at a special point called the end point or equivalent point. The reaction is called acid-base neutralization because the acid and base cancel each other forming a salt and water. The reaction can be represented:

A	+	B	--->	S	+	W
(acid)		(base)		(salt)		(water)

Before the acid and base are mixed there are only reactants. After they are mixed there are only products. The reaction following illustrates this.

	HCl	+	NaOH	--->	H_2O	+	NaCl
before mixing	1 mole		1 mole		0 moles		0 moles
after mixing	0 moles		0 moles		1 mole		1 mole

It is natural to assume that since the acid and base are used up in this reaction that the resulting solution will be neutral; that is, have pH=7. Let's study this idea. At the end point the only materials contributing to the pH of the solution are the products; that is, water and the salt. According to Investigation 14, the pH of pure water is 7. If the salt is neutral, then water is the only subtance effecting the pH of the mixture, which is then pH=7, indicating that the solution is neutral. NaCl is a salt which is neutral and therefore the acid-base neutralization between HCl and NaOH does result in a neutral solution.

However, not all salts are neutral. Consider the following acid-base neutralization:

$3HCl+Al(OH)_3 \rightarrow AlCl_3+3H_2$

In this reaction three moles of HCl are equivalent to one mole of $Al(OH)_3$. At the end point only H_2O and $AlCl_3$ are present. Water is neutral. However, during Investigation 12 you found that dissolving $AlCl_3$ in water gives an acid solution. Therefore, at the end point of this acid-base neutralization you are left with a solution which has a pH less than 7; that is, a solution which is not neutral.

There are three kinds of salts: acid-forming, base-forming, and neutral. You can determine what kind of salt you have if you know the strength of the acid and base you need to mix in order to produce the salt. In Tables 4 and 5 of Reading 13 there are lists of acids and bases. Strong acids like HCl, H_2SO_4, and HNO_3 have very large K_A values (that is they completely dissociate). Weak acids have K_A values less than one. A similar distinction exists between strong bases (NaOH, KOH) and weak bases.

The reaction between a strong acid and a strong base results in a neutral salt. For example:

$HCl+KOH \rightarrow H_2O+KCl$.

KCl is a neutral salt. The reaction between a strong acid and a weak base results in an acid-forming salt. For example:

$HCl+NH_4OH \rightarrow NH_4Cl+H_2O$.

NH_4Cl is an acid-forming salt. The reaction between a strong base and a weak acid is a base-forming salt, as shown:

$NaOH+HC_2H_3O_2 \rightarrow H_2O+NaC_2H_3O_2$.

This is the reaction you studied in Investigation 15:

$NaC_2H_3O_2$ is a base-forming salt.

INVESTIGATION 15: ACID-BASE INTERACTION

When a weak acid and a weak base are combined the situation is more complicated. You might have noticed in the examples above that if the acid was stronger than the base the resulting salt was an acid-forming salt. Some weak acids are stronger than others. Acetic acid is stronger than boric acid (see K_A values--the larger K_A value determines the stronger acid). The same is true of weak bases. If a weak acid is combined with a weak base, the stronger of the two determines the acid-base character of the salt. Which of the two is stronger can be determined by comparing the K_A and K_B values.

The following table summarizes this information.

Acid	Base	Resulting Salt
Strong	Strong	Neutral
Strong	Weak	Acid
Weak	Strong	Base
Weak	Weak	Depends on which is stronger.

INVESTIGATION 16: SALTS AND WATER

Name_____

Section_____

INTRODUCTION

You will be combining certain salts with water in this investigation. You already know, from your work in Investigation 15, that some salts are acid-forming, some are base-forming, and some are neutral. After filling in the blanks in Step 1, see if you can predict whether these salts are acid-forming, base-forming, or neutral. The goals of this investigation are to establish a reasonable explanation for the acid and base properties of salts and to look more deeply into the concept of acids and bases.

EXPLORATION

1. For each of the following salts, write the name and formula for the acid and base from which it could have been derived.

Salt	Acid (name and formula)	Base (name and formula)
Na_2CO_3		
NH_4Cl		
$NaC_2H_3O_2$		
Na_3PO_4		
$NH_4C_2H_3O_2$		

2. Test each of the salts from Part 1 by dissolving half a spoonful of the salt in approximately 25 ml of distilled water. Test each solution with litmus and for conductivity. Record your observations in the table that follows.

Salt (name and formula)	Litmus Test	Conductivity Test

INVESTIGATION 16: SALTS AND WATER

THE IDEA

1. List the salt solutions as acid, base, or neutral.

Acid	Base	Neutral

2. Offer a logical explanation for the formation of the acidic solution(s).

3. Offer a logical explanation for the formation of the basic solution(s).

INVESTIGATION 16: SALTS AND WATER

Name_____

Section_____

EXPANSION OF THE IDEA: Na_2HPO_4 and NaH_2PO_4

DEMONSTRATION

1. Write balanced equations for the results of the separate interactions of Na_2HPO_4 and NaH_2PO_4 with water.

2. As a result of the tests done on the mixture of Na_2HPO4 and NaH_2PO_4 solutions, describe the interactions that took place when the two solutions were mixed.

3. Record volumes (in <u>drops, ml,</u> etc.) required to cause significant changes in the pH for the various demonstrations.

Water and acid: _____

Water and base: _____

Na_2HPO_4 and acid: _____

Na_2HPO_4 and base: _____

NaH_2PO_4 and acid: _____

NaH_2PO_4 and base: _____

Mixture and acid: _____

Mixture and base: _____

4. Supply explanations for the tests performed.

INVESTIGATION 16: SALTS AND WATER

EXPANDING THE IDEA

READING 16: ALL DONATIONS ARE ACCEPTED

We have studied several situations in which observation and experiments forced the re-thinking of established theories and models. The discovery of isotopes, for example, initially threatened the existing atomic model and the periodic system of element classification. As time went on, the model for the atom was modified (though not rejected), but the periodic classification system remained intact.

Models and theories are of value only to the extent that they explain phenomena. If they fail to explain what happens, they must either be modified or rejected. Rejecting an established model outright is a hard thing for us to do because we are then in a position of having to construct a whole new system that will explain everything the old system explained, plus explaining things that the old system could not explain. It is much more desirable to attempt to resolve the problems by adjusting our existing model to explain the "new" phenomena while maintaining the explanations for the other phenomena.

Our initial laboratory work with acids and bases led to definitions something like the following.

ACID = A substance containing hydrogen, whose water solution contains an excess of H^+ ion. This excess of H^+ ion is caused by the ionization of the compound.

ACID = A substance containing hydroxide, whose water solution contains an excess of OH^- ion. This excess of OH^- ion is caused by the ionization of the compound.

Even before we wrote our first acid and base definitions, we had experienced some exceptions to the definitions--$AlCl_3$ forms an acid solution and CaO forms a base solution, for example. It is obvious that the compound $AlCl_3$ does not contain H^+, and CaO does not contain OH^-. It would seem, in light of our definitions of acids and bases, that water would have to be the substance responsible for the acidic and basic charcteristics of these solutions. There is a problem, however, because water is neutral, according to our tests on it, and water does not fit the acid or base definitions given above.

INVESTIGATION 16: SALTS AND WATER

We need a set of definitions that will include not only the hydrogen and hydroxide containing compounds, but compounds like $AlCl_3$ and CaO as well. Such definitions were proposed independently in 1923 by J. M. Bronsted of Denmark and T. M. Lowry of England. The definitions are as follows:

Acid = a proton donor.

Base = a proton acceptor.

Although these definitions look simple, their applications are far-reaching. We will spend the remainder of this reading discussing some significant ideas and applications relating to these definitions. Notice that the value of these definitions will be determined by how well they satisfy the criteria of our previous definitions while satisfying the need of including compounds not included by the first definition.

Both definitions contain the word <u>proton.</u> What does that mean? We have established that protons are in the nuclei of atoms and that the number of protons actually defines what the element is (atomic number). We know, further, that changing the number of protons in an atomic nucleus will change the atom from one element to another. In our discusssions of all chemical changes (including acids and bases), there have never been changes in the nuclei of the atoms involved. In other words, zinc atoms were zinc atoms whether they were in the elemental, metallic state or combined with other elements in compounds. This donation and acceptance of protons, therefore, must be such that no elements are changed. As you know, there is only one element in which the nucleus contains just one proton and that element is <u>hydrogen.</u> A neutral hydrogen atom consists of the proton (nucleus) and an electron outside the nucleus. Remove the electron and what remains is the H^+ ion, which is a proton! The nstedonsted-Lowry definitions, then, are dealing with transfers of H^+ ions (protons).

Notice that an acid, by this definition, <u>must contain hydrogen.</u> A base, by its definition, must be able to <u>attract and hold</u> (accept) an H^+ ion. We have seen that water can ionize to a small degree yielding an H^+ and an OH^- ion. The question arises as to what is causing water to ionize, and the only reasonable response is that other water has pulled the ionized water apart. The H^+ ion that is formed, therefore, is attracted and held to water. A simplified equation for this interaction is as follows.

$H_2O + H_2O \rightarrow H_3O^+ + OH^-$

Since the H^+ ion is attracted and held (accepted) by a water molecule, that water molecule is acting as a base (see the definition). The water molecule that is being ionized is acting as an acid (giving up, or donating, a proton). We see, then, that water can act as an acid or as a base, giving:

$H_2O + H_2O \rightarrow H_3O^+ + OH^-$
acid base

The strong attractions between the atoms in a water molecule would account for the quite small amount of ionization in water.

Once a substance has donated a proton, it will be capable of accepting back a proton. Also, once a substance has accepted a proton, it will be capable of donating it. In other words, when a substance acts like an acid (donates a proton), the remaining part of that substance can act like a base by exerting attraction on protons. When a substance picks up a proton, it becomes capable of acting like an acid (has a proton to donate). We can observe that in order for a substance to act as a Bronsted-Lowry acid there must also be a substance acting as a Bronsted-Lowry base. Furthermore, the end result of this acid-base interaction is that another acid and another base are produced. Our earlier equation now becomes:

$H_2O + H_2O \rightarrow H_3O^+ + OH^-$.
acid base acid base

Let's apply the ideas we have discussed to a situation involving an acid defined by our earlier definition--HCl. When HCl gas is put in water, the water ionizes the HCl:

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$

Notice the H_3O^+ (called the hydronium ion). We can see, from examination of its formula and from examination of the equations in which it is produced, that it is an H^+ ion attached to a water molecule. From our earlier discussion, we have:

$HCl + H_2O \rightarrow H_3O^+ + Cl^-$
acid base acid base

What is it that makes Cl^- a base? We see that H^+ was attached to Cl^- in the original compound, which shows that Cl^- must have (at one time or another)

accepted the H^+ ion. Furthermore, the Cl^- has a negative electrical charge which will attract any positive electrical charge that comes near it. You might reason that any negatively charged ion could be a Bronsted-Lowry base. Although this idea would certainly be theoretically valid, there are times when practical situations give us some surprises. Notice the last equation, for example. In the equation, the water (a neutral molecule) appears to exert a greater attraction force on the H^+ ion than does the chloride ion (with a negative charge).

Well, then, is the Cl^- ion a base? Yes! The important point here is that the real situation (what happens when substances are combined) will determine which substance will be most effective as an acid or as a base.

Now that we have seen how the Bronsted-Lowry definitions can be used for substances defined by our first definitions, let's see if they will apply to other substances not covered by the first definitions. Consider the reaction of CaO with water. We have seen that there must be an acid and a base (donor and acceptor) in a system, which leads to:

$$\underset{}{Ca^{2+}} + \underset{\text{base}}{O^{2-}} + \underset{\text{acid}}{H_2O} \longrightarrow \underset{}{Ca^{2+}} + \underset{\text{acid}}{OH^-} + \underset{\text{base}}{OH^-}$$

There are two interesting aspects of this reaction that can be discussed. First, notice that the calcium ion is neither an acid nor a base. It would not be possible to consider it an acid because it does not have an H^+ ion to donate. It could not be a base because the positive charge of the calcium ion would repel, not attract (or accept) a proton. An ion such as this one is called a spectator ion since it doesn't function as an acid or as a base. The second interesting aspect of this reaction is that the acid and the base on the product side of the reaction are both hydroxide ions!

You might be wondering about the sense of a system that would introduce something as confusing as two hydroxide ions each having a different function. The reality of the situation is not that the two ions would have different functions, but the two ions have different designations. Recall the earlier discussion where the statement was made that a substance, in acting like an acid, becomes a base (and vice versa). This situation always exists in discussions of Bronsted-Lowry acids and bases. The oxide ion, therefore, acts as a base by accepting a proton, but now it has a proton that it could donate and is designated as an acid. Similarly, water donates a proton (acid action) and the hydroxide ion that is left could accept a proton and

is designated as a base. The acid-base pairs (that is, O^{2-} to OH^- and H_2O to OH^-) are called conjugates. A substance, in acting like an acid, will produce a conjugate base, and a substance that acts like a base will produce a conjugate acid.

STRONG AND WEAK

The strength of acids and bases was discussed in some detail in READING 13b: THOSE MIGHTY ACIDS, but it would be good for us to look at some aspects of acid and base strength related to the Bronsted-Lowry acid-base theory. You have obtained observational evidence that clearly shows hydrochloric acid to be a stronger acid than boric acid. Possibly the difference in the effect on litmus and certainly the difference in conductivity tests indicated that HCl ionized much more thoroughly than H_3BO_3. From our early definitions we would expect a substance to be more acidic if the number of H^+ ions were large. Let's look at an explanation of these strong and weak characteristics using the Bronsted-Lowry framework.

The amount of ionization will be controlled by how strongly a substance is held together and how strong the force is that is trying to pull it apart. In our two examples (HCl and H_3BO_3), water is the substance that is attempting to pull the acids apart into ions. If we can assume that the pulling force from the water is the same on both acids, then any difference in the amount they ionize will be determined by how strongly the molecules of the acids are held together. It follows that the attraction between hydrogen and chloride is much less than the attraction between hydrogen and the dihydrogen borate ion ($H_2BO_3^-$).

Stating this in Bronsted-Lowry terms, HCl donates protons more readily than does H_3BO_3.

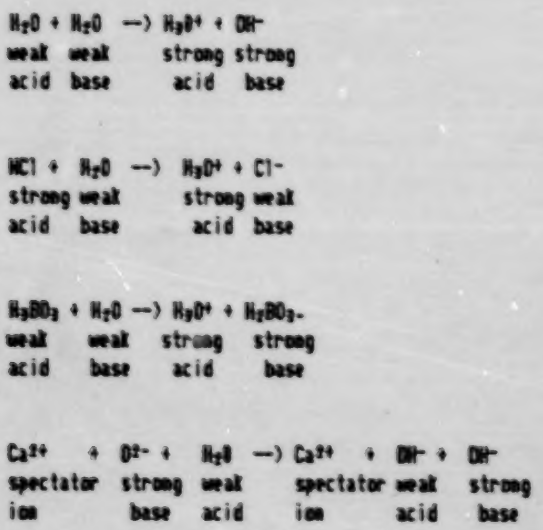
Now let's examine the conjugate bases produced by the ionization (proton donation) of these two acids. Since a base is defined in terms of attracting protons, it would seem to be reasonable that a substance with a strong attraction for protons would do a more thorough job of acting like a base than a substance that had a weaker attraction. The reason that a substance would be reluctant to donate a proton is that it is eager to accept a proton. So, if we say that boric acid is a weak acid, then its conjugate base ($H_2BO_3^-$) must be a strong base. Conversely, since HCl is a strong acid, the chloride ion (conjugate base) must be a weak base.

Water is also participating in these reactions. It is acting as a base in both situations, and the question arises as to whether water is a strong base or a weak base. Consider the conjugate pair discussed below.

H_3BO_3	and	$H_2BO_3^-$
weak acid		strong base

Recall that there is very little ionization of boric acid in water. Similarly, pure water does not ionize very much. We might say, therefore, that water is a weak acid. The base strength of water, however, could best be determined in a different way. Since bases accept H^+ ions, the strength of the base would generally increase as the amount of attraction for H^+ ions increases. Therefore, we would <u>generally</u> expect the attraction of a negtive ion for H^+ ions to be greater than the attraction of neutral molecules for H^+ ions (there is an apparent exception when the Cl^- ion and water are compared). By comparison to other substances that can act as bases, we could then consider water to be a weak base.

Using the conjugate strength arguments from before, we would have to conclude that water's conjugate--the hydronium ion--is a strong acid. Notice the designations that result from this discussion when they are applied to some of the systems we have discussed.


H2O + H2O → H3O+ + OH-
weak weak strong strong
acid base acid base
HCl + H2O → H3O+ + Cl-
strong weak strong weak
acid base acid base
H3BO3 + H2O → H3O+ + H2BO3-
weak weak strong strong
acid base acid base
Ca2+ + O2- + H2O → Ca2+ + OH- + OH-
spectator strong weak spectator weak strong
ion base acid ion acid base

INVESTIGATION 16: SALTS AND WATER

Name_____

Section_____

READING 16: QUESTIONS

1. In Investigation 5 you added water to magnesium nitride, which had formed in your system, and got a reaction that can be described by the following equation:

   $Mg_3N_2 + 6H_2O \rightarrow 2NH_3 + 3Mg(OH)_2$

   a. List the spectator ion(s) in this reaction.

   b. How many single proton transfer reactions would be involved in making this reaction occur?

   c. Write a balanced equation for each of the single proton transfer equations. (NOTE: The sum of these equations must equal the overall balanced equation above.)

2. Ammonium carbonate will decompose at atmospheric pressure to produce ammonia gas (NH_3). Write a balanced equation for this decomposition, label the reactant acid and base and the conjugate acid and base.

3. Label as Bronsted-Lowry acids and bases the appropriate parts of the equation showing the reaction of barium hydroxide with sulfuric acid.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

Name_____

Section_____

INTRODUCTION

During INVESTIGATION 13: ARRHENIUS SOLUTIONS, you developed ideas that allowed you to recognize and specify whether a substance was an acid or a base. In this investigation you will be studying one compound: methanol (CH_3OH). In Steps 1 and 2 you will predict some characteristics based on your previous study, and then you will test these predictions in the laboratory in Steps 3 and 4.

EXPLORATION

1. State how methanol will affect litmus paper and whether or not it will conduct electricity. Give supporting evidence for your predictions.

2. State how a solution of water and methanol will affect litmus paper and whether or not it will conduct electricity. Will the concentration of solution affect these properties? Give supporting evidence for your prediction.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

3. Test methanol for conductivity and acid-base characteristics. Describe your tests and record the results of your tests.

4. Test solutions of methanol and water for conductivity and acid-base characteristics. Describe your tests (including the relative proportions of the two liquids) and record the results of your tests.

THE IDEA

1. Did one or more of the solutions of methanol and water ionize? If so, what ions were formed? If not, why not?

2. Offer an explanation for the results of the tests in light of the fact that methanol <u>is</u> soluble in water.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET a: ELECTRON-DOT STRUCTURES

1. Write electron-dot structures for the first twenty elements.

hydrogen

helium

lithium

beryllium

boron

carbon

nitrogen

oxygen

fluorine	neon
sodium	magnesium
aluminum	silicon
phosphorus	sulfur
chlorine	argon
potassium	calcium

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET b: ELECTRON-DOT STRUCTURES

2. Write electron-dot structures for the following molecules.

methane (CH_4)

ammonia (NH_3)

water (H_2O)

carbon dioxide (CO_2)

hydrogen peroxide (H_2O_2)

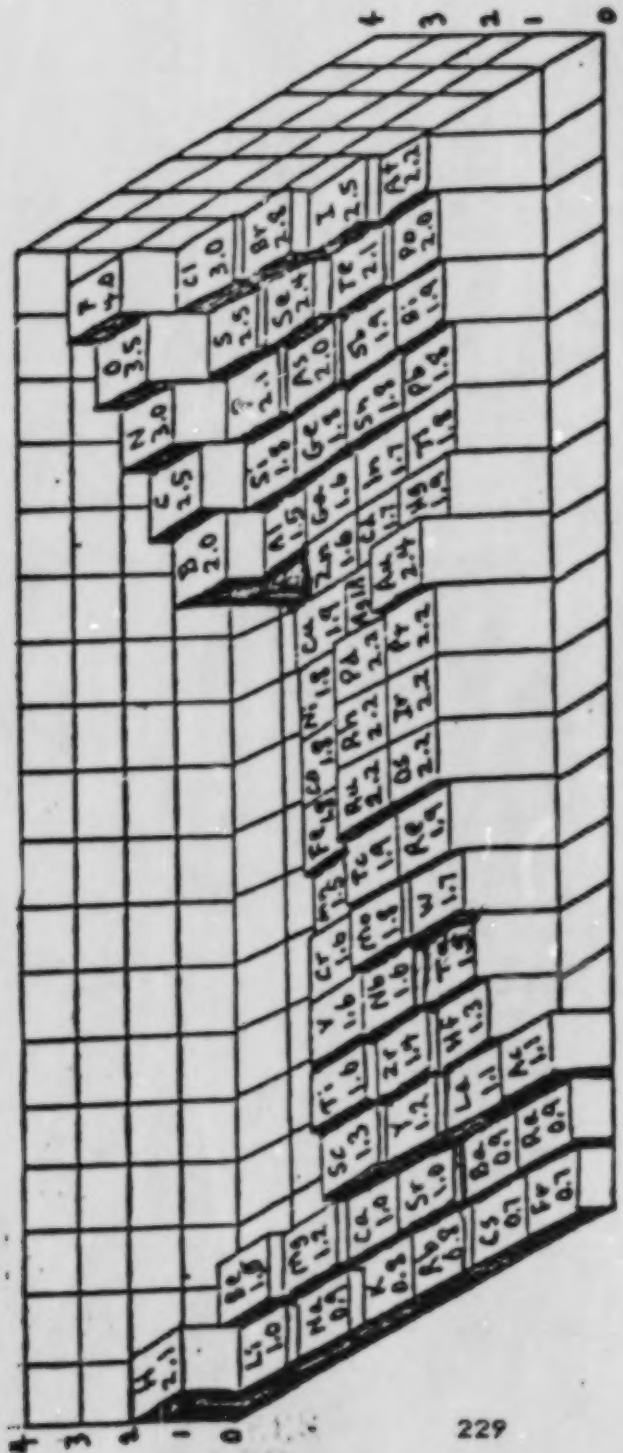
hydrogen (H_2)

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

nitrogen (N_2)	oxygen (O_2)
fluorine (F_2)	chlorine (Cl_2)
bromine (Br_2)	iodine (I_2)

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

ELECTRONEGATIVITIES OF THE ELEMENTS



INVESTIGATION 17: CHARACTERISTICS OF METHANOL

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET c: ELECTRONEGATIVITY EXERCISES

For each of the following molecules, determine the electronegativity difference between the elements and state which of the elements has the greater attraction for the electrons in the bonds between the atoms. The first one is done as an example.

molecule	formula	electronegativity difference	element with the greatest attraction for bonding electrons
carbon monoxide	CO	3.5-2.5=1	oxygen
hydrogen			
nitrogen			
bromine			
hydrogen iodide			
hydrogen bromide			
hydrogen chloride			
aluminum nitride			
cesium bromide			
lithium fluoride			
potassium chloride			

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

molecule	formula	electronegativity difference	element with the greatest attraction for bonding electrons
magnesium oxide			
barium oxide			
iron (II) oxide			
silver fluoride			

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET d: ELECTRON AND ATOM ARRANGEMENTS

1. Make drawings that will show the arrangement of the electrons in the valence level for each of the following ions.

   EXAMPLE: N^{3-}

   (a) O^{2-}

   (b) Br^-

2. Describe the shape represented by the drawings in Question 1.

3. Draw diagrams to indicate the shapes of the following molecules.

   EXAMPLE: methane (CH_4)

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

(a) water

(b) carbon tetrachloride (CCl_4)

(c) ammonia (NH_3)

(d) methanol (CH_3OH)

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

EXPANSION OF THE IDEA

READING a: ARE ELECTRONS SALTY?

We have studied two basic types of chemical bonding so far--ionic and covalent. There are characteristic physical properties of substances that are bonded ionically, properties of which can often be used to distinguish these substances from substances that are made up of atoms held together with covalent bonds. A familiar ionic substance is sodium chloride and a familiar covalent substance is wax. Let's compare some physical properties of these two substances.

salt	solid wax
1. hard	1. soft
2. brittle	2. can be slightly deformed without breaking
3. definite crystal structure	3. no definite crystal structure
4. conductivity a. heat--poor b. electricity--poor	4. conductivity a. heat--poor b. electricity--poor
5. high melting point	5. low melting point

Notice that in order to observe physical properties of substances, we must have large collections of molecules of those substances. Also, the definition of physical properties does not involve chemical changes within the substance. When we investigate physical properties, therefore, we are looking at <u>intermolecular</u> forces (between molecules) in covalent substances and <u>interionic</u> forces (between ions) in ionic substances. A glance at the lists of properties above will confirm the fact that the intermolecular attractions between covalent molecules are generally much weaker than the interionic attractions in ionic compounds. Ionic and covalent substances, furthermore, are almost always compounds of two or more elements.

The reason for all this review is to give us a base of information from which we can look at metals (a third common type of solid substance) and relate their physical properties to the way in which they are held together. A list of some physical properties of metals follows.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

METALS

1. Hard

2. Malleable--this means that metals can be beaten into shapes without breaking

3. Definite crystal structure

4. Conductivity
   a. heat--good
   b. electricity--good

5. High melting point

NOTE: These are <u>general</u> characteristics, and there will be a few exceptions to most of these properties; for example, mercury has a low melting point.

We can expect, from the comparison of the metal properties with the general properties of ionic and covalent substances, that metals are neither held together with ionic bonds nor with covalent bonds. Although some metal properties correspond to those of ionic and/or covalent substances, there are several major differences that occur between the properties of metals and the other two types of substances. Our observations show that the type of bonding within a substance (INTRAmolecular attraction) has an effect on the INTERmolecular attractions, and we have seen that the intermolecular attractions affect the physical properties of the substance. We will conclude, therefore, that the bonding of metal atoms to each other is a unique type of bonding.

Our problem now is to try to build a model for the bonding in metals that will explain the physical properties of these substances. One very prominent property of metals is their ability to conduct electricity in the solid state. This property is not a characteristic of either ionic or covalent substances (in the solid state). We will work from this physical property to begin developing a model for metallic bonding.

During the discussion of electrical charge movements and interactions in the electroscope, we found that our atomic model required that the electrons be the charges that moved for two reasons.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

1. Electrons are lightweight (more easily moved than the much heavier positive charges) and accessible (our model shows them to be on the outer portions of the atom).

2. Movement of positive charge necessitates breaking up the nucleus or moving an entire nucleus, both of which are contrary to our observations.

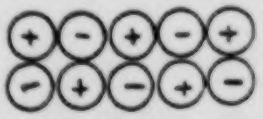
Our bonding model must allow some electrons in a metal to move from one atom to another with ease. Of course, this could only occur in the solid or liquid states because in the gaseous state--the state elements are in when ionization energy data are gathered--conduction (current flow) occurs only after much more energy is added than would be necessary in a collected (or aggregate) state.

It is reasonable to assume that the valence electrons in the metals would be involved in the bonding, but anything that would restrict valence electron movement--as in the ionic and covalent bonding models--would not satisfy our needs. We could assume, therefore, that the valence electrons of the atoms are contributed freely to other atoms in a crystal of metal. Notice that this would be a sharing effect (covalent in nature), but the sharing would not take place in one direction only as in covalent bonding. Instead, the sharing would be among many atoms, and the valence electrons would have to be allowed to move freely among the atoms to achieve this sharing effect. Our model, then, could be characterized as a collection of positive metal ions bathed in a sea of electrons. We see that a given metal atom no longer owns its own valence electrons, but there are enough electrons in the crystal to maintain a neutral overall charge. There would also be a strong overall attraction among the atoms because of the ion-like attraction of the positive metal ions on the electron sea. This model appears to be a combination of the models for ionic and covalent bonding, but it also has characteristics that are different from both of the other models.

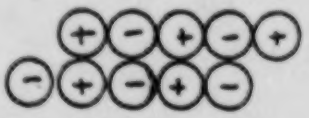
The freedom of electron movement in the sea would explain why a small amount of electrical force would cause a directed movement of charge (electrons), which is an electrical current. The strong attraction among the metal atoms, furthermore, would account for the strength and hardness of metals. Let's see how this bonding model can be applied to another physical property of metals--their malleability.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

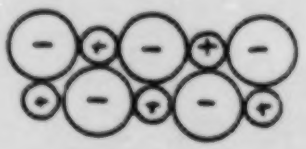
It has been stated that metal atoms arrange themselves in orderly crystalline structures in the solid state. You will recall from the list of properties for NaCl given earlier that ionic substances also characteristically solidify into orderly crystalline structures. Two major differences between these crystal types will concern us here: (1) the difference in bonding type, and (2) the sizes of the particles that make up the crystals. In an ionic crystal there is the strong attraction of oppositely-charged ions that serves to hold the crystal together. This arrangement can be represented by:



If the top row of ions (in the figure above) were to be displaced one ion to the right (a situation that could arise from the crystal being hit with a strong force), like-charged ions would be near each other and repulsion would occur between the row of ions.



This repulsion will account for the tendency of a crystal to break apart when it is hammered. Many ionic crystals are made up of ions whose sizes are quite different. If the negative ion, for example, were larger than the positive ion, we might expect a picture like the following.



Such an arrangement would make the displacement of particles (past each other) even more difficult than it would have been when the ions were the same size. Both situations, however, would contribute to the brittleness of ionic crystals.

Metal crystals, on the other hand, are made up of atoms that are all the same size (assuming that the crystal contains only one metal). Additionally, the non-directional character of the bonds among the atoms (sea of electrons) would allow for strong attraction between them with little regard for their relative positions. A deforming force, therefore, can shift the positions of some of the atoms in the crystal without destroying the attractions among the atoms, as it would in an ionic crystal. The atoms being all of the same size helps this property of malleability by allowing the atoms to slip by one another with minimum interference. The existence of differently-sized atoms in a crystal of an alloy can be used to partially explain the general tendency of alloys to be harder than a pure metal.

This basic model for bonding of metal atoms to each other is useful in helping to explain several other characteristic properties of metals as well. Some such properties are good heat conductivity, large temperature differences between melting point and boiling point, and metallic luster.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

Name_____

Section_____

READING a: QUESTIONS

1. Using the model for metallic bonding developed in this reading, offer an explanation for the good heat conductivity characteristics of metals.

2. As temperature increases, the electrical conductivity of metals decreases. Offer an explanation for this phenomenon in terms of the metallic bonding model.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

EXPANSION OF THE IDEA

READING b: ARE ELECTRONS COORDINATED?

You have had the opportunity to deduce and write the formulas for <u>many</u> compounds by combining positive and negative ions. Some of these ions have been simple, monatomic ions, while others have been polyatomic (or <u>complex</u>) ions. The monatomic A-column ions have been explained in terms of the formation of noble gas electron configurations, but you were asked to postpone discussion of the polyatomic ions until <u>later.</u> Well, it's <u>later,</u> and even now we won't exhaust the topic--by any means--but we will look at some polyatomic ions and try to find some useful ideas from our search.

As examples of complex ions, let's focus our attention on the ammonium ion--NH_4^+, the hydronium ion--H_3O^+, and the sulfate ion--$SO_4^=$.

The first thing that we will state about each of these ions is that they are held together with--you guessed it!--<u>covalent</u> bonds. In the covalent combinations, however, here turns out to be either a deficiency (positive ions) or an excess (negative ions) of electrons for the number of atoms present. As long as the complex ion is reasonably small, the net ionic charge has the effect of making the entire particle exhibit the ionic charge, as though it were a single atom with a net charge. It will, therefore, repel like charges from any direction and attract opposite charges from any direction.

Given the foregoing information, consider the building of an ammonium ion from its elements. If the ion is covalently bonded, then we should be able to gain some information on its structure by working with electron-dot structures for the elements involved. Apparently, from the formula (NH_4^+), we will need one nitrogen atom and four hydrogen atoms, whose electron-dot structures are as follows.

$\cdot\ddot{\underset{\cdot}{N}}\cdot$ and $H^{\bullet}$ (four of these)

There is no problem in constructing this ion to the point of the ammonia molecule, but then a problem occurs in that there is no place for the fourth hydrogen atom to bond.

$$H:\ddot{\underset{\ddot{H}}{N}}:H \qquad \text{and } H^{\bullet} \text{ (one of these)}$$

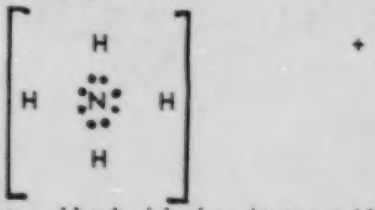
It is clear that, even if the fourth hydrogen could bond, the molecule would be neutral and not contain a net charge.

You have probably already anticipated the situation dictated by the ion formula and our present condition. The fourth hydrogen will have to be an H^+ ion, not a neutral atom! Where does the H^+ ion come from? Well, it will have to come from some substance that is capable of giving up a proton--a Bronsted-Lowry acid. That would make ammonia a base (accepting the H^+ ion). We can see that water is one possible source of this H^+ ion from the following equation.

$$\underset{\text{base}}{NH_3} + \underset{\text{acid}}{H_2O} \longrightarrow \underset{\text{acid}}{NH_4^+} + \underset{\text{base}}{OH^-}$$

Let's look more closely now at the interaction of the H^+ ion with the ammonia molecule.

Since the H^+ ion will only be attracted to negative electrical charge, the place at which it will be attracted to the ammonia molecule is clearly defined. It can only be attracted at the place where the non-binding pair of electrons is on the nitrogen atom. Now, observe the resulting picture of the ion.



You will notice that it is impossible to tell which of the hydrogen atoms in this ion is the hydrogen ion; in other words, all of the bonds between nitrogen and the hydrogen atoms are <u>equivalent.</u> "But wait!", you say. "The H^+ ion is the one at the top of the drawing!" Recall that the non-bonding pair of electrons does not have to occupy any particular place in the drawing. It could be on top, bottom, left side, or right side of the element symbol, as <u>convenience</u> would dictate. Notice, also, that the positive effect is spread over the entire ion.

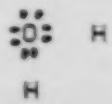
It is true that we cannot tell which of the hydrogens was originally the ion. All four bonds between hydrogen and nitrogen are covalent and equivalent. The difference is the way in which one of the covalent bonds are formed. The first three bonds are formed by the sharing of one electron from nitrogen and one electron from hydrogen. The fourth bond formed when the ion, which had room for two electrons,

shared the two electrons that were both contributed to the bond by the nitrogen atom. Such a bond is called a coordinate covalent bond. It is no less a covalent bond than are the ones formed by atoms which each contribute electrons to the bond formation (see the ion diagram above). In a coordinate covalent bond both bonding electrons come from one atom.

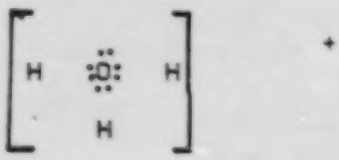
The hydronium ion is another example of coordinate covalent bonding. We have seen this ion form in several reactions such as:

$$\underset{\text{acid}}{HCl} + \underset{\text{base}}{H_2O} \longrightarrow \underset{\text{acid}}{H_3O^+} + \underset{\text{base}}{Cl^-}$$

Notice that the water molecule has two non-bonding pairs of electrons according to the following diagram.

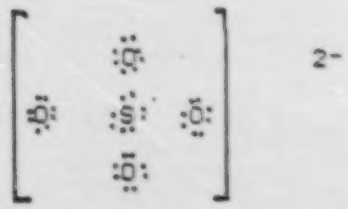


The H^+ ion forms a coordinate covalent bond with one of these non-bonding pairs of electrons to produce the hydronium ion.



You may be wondering why H_4O^{2+} does not form. The several possible explanations for H_4O^{2+} not forming will be left as an exercise for the student; but, as always, we must reason from observation to model and not the other way around. In other words, since H_4O^{2+} does not occur, what explanations can be put forth to support the observed phenomenon?

Reviewing what has been said about coordinate covalent bonds, we find that there must be one chemical species (that is, atom, molecule, or ion) that has a pair of electrons that it can donate to the bond. There must also be a chemical species that has room for, or can accept, a pair of electrons. Consider the construction of a sulfate ion. The following structure will account for both its formula and the charge. Carefully inspect the structure to see that it is reasonable.



We observe that the ion is made up of two elements, both of which are VI-A column elements, so they each have six valence electrons (when they are neutral). There are several ways that the structure above could be synthesized, but each method would ultimately require the formation of coordinate covalent bonds to the sulfur in order to achieve the stated structure. Other construction methods would not require as many coordinate covalent bonds, but each structure would contain some.

There are structures for some polyatomic ions that cannot be explained in terms of coordinate covalent bonds alone. The explanations of these complex ions (nitrate is an example) is beyond the scope of this reading. The reason for mentioning the existence of such ions is to alert you to the possibility that you may come across an ion whose structure cannot be explained according to the arguments in this reading.

There is an important type of complex ion that results from the coordinate covalent bonding of several neutral molecules (or sometimes ions) to a single, simple ion. Pure, dry copper (II) sulfate is a white powder, but it dissolves in water to give a blue solution. If the water is evaporated slowly from this solution, blue crystals are formed. Analysis, by weight of the solid before compared to the solid after the dissolving procedure, shows the addition of weight corresponding to five weights of water for each $CuSO_4$ molecule. Studies of structure have shown that four of the water molecules have donated a pair of electrons each to the Cu^{2+} ion forming what is known as a coordination complex. The formula for this complex ion, formed by four coordinate covalent bonds, can be written as

$$[Cu(H_2O)_4]^{2+}.$$

This coordination complex is what imparts the blue color to a water solution of this compound.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

If ammonia is added to a solution containing the $[Cu(H_2O)_4]^{2+}$ complex, the color of the solution becomes a much darker blue. The reason for the color change is a chemical reaction in which the coordinating water is displaced by the ammonia molecules to yield $[Cu(NH_3)_4]^{2+}$.

Another example of coordination complexes refers to an insoluble compound you have produced--AgCl. As you know, this substance is totally insoluble in water for all practical purposes. If a concentrated ammonia solution is added, however, the silver chloride will dissolve. Apparently the ammonia, unlike the water, is able to force its way between the silver and chloride ions and form a complex ion

$$[Ag(NH_3)_2]^+$$

that water <u>can</u> dissolve.

INVESTIGATION 17: CHARACTERISTICS OF METHANOL

Name_____

Section_____

READING b: QUESTIONS

1. Propose a structure for the perchlorate ion--ClO_4^---and explain how it could form.

2. Suggest reasons why H_4O^{2+} does not form.

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

Name_____

Section_____

INTRODUCTION

As the title of this investigation suggests, you will be studying the relationship between the structure of various substances and some of their physical properties--specifically, solubility, melting point, and boiling point. This study will allow you an opportunity to utilize your learning about covalent bonding.

EXPLORATION

1. Use approximately 2 ml of the liquid and a small amount of solid. When combining two liquids, use equal volumes of the two liquids. Stopper and shake the test tubes vigorously.

CAUTION: If you get any of the solids or liquids (other than water) on your hands, wash your hands with soap and water.

2. Label the various test tubes for the tests and keep them and their contents to compare with the other tests done. Such comparisons will help you to decide about the solubility of various substances.

3. Describe the results of your solubility tests in the appropriate spaces.

4. When you have completed all the tests, dispose of the contents of the test tubes according to the instructions of your teacher.

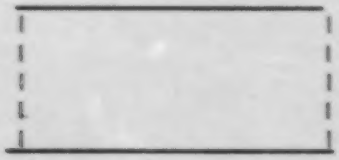
INVESTIGATION 18: SOLUBILITY AND STRUCTURE

SOLUBILITY DATA

	Potassium Chloride	Iodine	Distilled Water
distilled water			
hexane C_6H_{14}			
carbon tetrachloride CCl_4			

carbon tetrachloride

hexane



INVESTIGATION 18: SOLUBILITY AND STRUCTURE

THE IDEA

1. Classify the substances used in this investigation as polar or non-polar.

POLAR	NON-POLAR

2. Explain how you arrived at the classification you gave for:

   a. hexane

   b. water

3. What pattern in solubility can be deduced from this investigation?

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

Name_____

Section_____

EXPANSION OF THE IDEA

SOLUBILITY WITH ALCOHOLS: EXPERIMENT

Follow the same procedures that were used in EXPLORATION to test the solubility of the combinations of substances in the table below.

	distilled water	hexane (C_6H_{14})	carbon tetrachloride (CCl_4)
methanol CH_3OH			
ethanol C_2H_5OH			

	potassium chloride	iodine
methanol		
ethanol		

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

QUESTIONS

1. Are methanol and ethanol polar or non-polar? Explain your answer.

2. How do the data obtained in these tests compare to those obtained in the EXPLORATION?

3. Offer an explanation for any deviations from the solubility rules developed during your class discussion.

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

Name_____

Section_____

EXPANSION OF THE IDEA

PROBLEM SET: MELTING POINTS, BOILING POINTS, AND STRUCTURE

Information on four groups of compounds is given in the table below. After filling in the blank spaces in the table, use the information to answer the questions that follow.

	FORMULA	MOL.WT.	M.P.(°C)	B.P.(°C)	POLAR/ NON-POLAR
a-1 methane	CH_4		-183	-162	
a-2 silicon hydride	SiH_4		-185	-112	
a-3 germanium hydride	GeH_4?		-165	- 88.5	
a-4 tin hydride	SnH_4		-150	- 52	
b-1 ammonia	NH_3		- 78	- 33	
b-2 phosphorus hydride	PH_3		-133	- 87	
b-3 arsenic hydride	AsH_3		-116	- 55	
b-4 antimony hydride	SbH_3		- 88	- 17	
c-1 hydrogen oxide			0	100	
c-2 hydrogen sulfide			- 85.5	- 61	
c-3 hydrogen selenide			- 60	- 42	
c-4 hydrogen telluride			- 49	- 2	

	FORMULA	MOL. WT.	M.P.(°C)	B.P.(°C)	NON-POLAR
d-1 hydrogen fluoride			- 83	19.5	
d-2 hydrogen chloride			-115	- 85	
d-3 hydrogen bromide			- 88	- 67	
d-4 hydrogen iodide			- 51	- 35	

1. For each of the groups of compounds in the table, list similarities, differences, and any trends that you can observe among the characteristics listed.

	SIMILARITIES	DIFFERENCES	TRENDS
group a			
group b			
group c			
group d			

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

2. What relationship exists between the molecular weights and the melting points and boiling points of these substances?

3. Offer a logical reason for the relationship between the molecular weights and the melting points and boiling points.

4. Offer an explanation for the exceptions to the relationship stated in Question 2; that is, for compounds b-1, c-1, and d-1. (HINT: Take electronegativity differences into account.)

5. Consider the following compounds: NaF, NaCl, NaBr, and NaI. Describe these compounds in terms of their polarity.

6. How would the melting points and boiling points of the compounds in Question 5 compare to the same properties of compounds in groups a through d which have similar molecular weights? Explain the difference.

7. If you used a reference source to find the melting and boiling points for these substances, you found that for both properties (M.P. and B.P.) the values compare in the following way:

NaF > NaCl > NaBr > NaI

Offer an explanation for this phenomenon.

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

EXPANSION OF THE IDEA

READING a: BETWEEN THE MOLECULES (An Attractive Story)

A molecule has been defined as two or more atoms that are chemically combined. Examples satisfying this definition would include things such as the seven diatomic elements, water and methanol. Common to all of these examples is the type of attraction that holds the atoms together in the respective molecules--covalent bonding. That attraction could also be called intramolecular attraction, since intra- is a prefix meaning within. These examples (and all other covalent molecules, for that matter) also exert intermolecular attractions on other molecules that come near them. Note that the prefix inter- means between or among. These intermolecular attractions are of neither the same character nor of the same strength as the intramolecular, bonding attractions. They are, in fact, much weaker than bonding attractions. That the boiling of a liquid is a physical change and not a chemical change is evidence that intermolecular attractions are weaker than intramolecular attractions. What is the character of these intermolecular attractions? We will analyze some of the ways molecules could interact with each other to establish models for these attractions between (or among) molecules.

Consider, first of all, the elemental halogen molecules--F_2, Cl_2, Br_2 and I_2.

As you know, these molecules are covalently bonded and the bonds are non-polar. The molecular weights and boiling points for these elements are as follows:

Element	Molecular Weight	Boiling Point (°C)
F_2	38	-188
Cl_2	71	- 35
Br_2	160	+ 59
I_2	254	+184

The data clearly show that there are intermolecular forces of attraction in each of these substances. If that were not so, it would not take a measurable amount of energy to force the molecules to separate (boil). Furthermore, the amount of attraction between the molecules increases with increasing molecular weight.

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

Since these molecules are non-polar, it would be reasonable to ask how they could even attract each other. We need to come up with a way that these neutral, non-polar molecules can attract each other without forming a chemical bond. We have assumed that all matter is in constant motion. Our atomic model has suggested, furthermore, that the motion of electrons is great by comparison to that of the nuclei of the atoms. The very fact of constant, rapid electron motion could cause the development of what could be called an instantaneous dipole. Compare the following instantaneous charge distribution models for a fluorine molecule.

F:F F:F F:F

(a) (b) (c)

Figures (b) and (c) would represent instants when the electron cloud had shifted to create dipoles on the molecule, while figure (a) would represent an even distribution of negative charge throughout the molecule. Notice that there would be a great number of instantaneous dipole arrangements that could be drawn (with widely varying dipolar strengths), but there is only one arrangement that can be drawn to indicate no dipole. Of course, a time average of all the instantaneous arrangements woud have to yield figure (a). What about the effect of these instantaneous dipoles among a group of molecules? We would expect that part of the time these instantaneous dipoles would cause attraction between molecules and part of the time these dipoles would cause repulsion. The net effect would have to be attraction, which would simply mean that the attractive situations outnumber and overpower the repulsive situations. These weak intermolecular attractions are known as dispersion forces.

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

Now that we have the molecules being attracted to each other by the random movements of electrons, it would be nice to have an explanation for the apparent increase in attraction with increasing molecular weight. Observe that as molecular weight increases, it does so because of the greater number of protons and neutrons in the nuclei of the atoms in the molecule. With an increase in the number of protons there will also be--right again!--an increase in the number of electrons. The greater the number of electrons, the more spread out (or diffuse) they will be. The more diffuse the electron cloud, the more spongy it is--the easier it would be to cause a shifting of the charge. These electrons are not held as rigidly by the nucleus as ones that are closer (as in F_2), so the instantaneous dipoles could form more freely and in greater magnitude. This effect would be magnified by the greater surface of interaction between the larger molecules.

In polar molecules, the dipoles are permanent. Instantaneous dipoles (giving rise to dispersion forces) would have to be present, but it would seem logical that the net attraction from a continuous (or permanent) dipole would be greater. For polar substances, then, we would expect to get dipole to dipole and dispersion attractions. If this reasoning is profitable, it should predict that a polar molecule, of the same (or similar) weight as a non-polar molecule, should have a higher boiling point than the non-polar molecule. Do data support this idea?

A third type of intermolecular attraction that could be assumed could arise as the result of a polar molecule causing a shifting of charge in a non-polar molecule (or in a non-polar part of a large polar molecule). This action is called inducing a dipole into the normally non-polar particle. The induced dipole would not have to remain, but attractions could result while the shift in charge remained. The weakness of these attractions can be illustrated by the poor solubility of hexane, carbon tetrachloride, and other non-polar liquids in water.

This discussion has led to the definition of three types of attractive, intermolecular forces: dispersion forces, dipole-dipole forces, and dipole-induced dipole forces. These weak attractive forces between molecules are sometimes referred to as van der Waals forces, after the Dutch physicist who first studied the importance of intermolecular attractive forces.

INVESTIGATION 18: SOLUBILITY AND STRUCTURE

EXPANSION OF THE IDEA

READING b: HYDROGEN BONDING

As the model for attraction between molecules of substances is applied, we find that the explanations are generally valid and complete. Bigger, heavier molecules generally have greater intermolecular forces of attraction than do smaller, lighter molecules. Polar molecules generally have stronger intermolecular attractions than non-polar molecules of similar size or weight. There are, however, some distinct exceptions to the data predicted by the forces discussed in Reading a. For example, look carefully at the following table of boiling point data for compounds containing hydrogen and V-A, VI-A, and VII-A elements, respectively.

BOILING POINTS (°C)

NH_3 = -33.4	H_2O = 100	HF = 19.4
PH_3 = -87.4	H_2S = -62	HCl = -83.7
AsH_3 = -55	H_2Se = -42	HBr = -67
SbH_3 = -17	H_2Te = -4	HI = -35.4

Notice that the last three compounds in each series have boiling points that follow the aforestated model of intermolecular attractions very nicely. In order for the three top-listed compounds (H_3N, H_2O, and HF) to follow the patterns already established, they should each have the <u>lowest</u> boiling point of the compounds listed in their respective series! They are, after all, the lightest compound of their series.

Of course, the data show that there must be considerably more attraction between molecules of the top-listed compounds than for any others in their respective series (except for H_3Sb, when compared to H_3N). Examination of the formulas within a series shows that the molecular shapes of all the substancs in a series would be essentially the same. The magnitude of the polarity of these molecules, however, will decrease (down the series) because of the decrease in electronegativity of the non-hydrogen element of the compound. This will lower the dipole-dipole attractions. However, if these data could be accounted for in terms of increased polarity alone, the values of the data should slowly decrease down a series instead of showing a sharp decrease between the first and second compounds and a gradual increase after that.

It is clear that the top-listed compounds will be the most polar for their series, but that high polarity gives rise to an unexpectedly high attractive effect. Now in each of these cases, hydrogen is covalently bonded to a small, highly electronegative atom. (Recall that electronegativity deals with the attraction that an atom has for a bonding pair of electrons.) The bonding pair of electrons defines the only electrons associated with the hydrogen atoms in these molecules. If these bonding electrons are shifted strongly toward the other element because of the second element's high electronegativity, the positive charge of the hydrogen nucleus becomes exposed. This strong partial positive charge will, most certainly, attract negative charge from whatever source. The source of negative charge here will be the partially negative part of another molecule of the substance. This may sound a lot like the dipole-dipole attraction discussed earlier, and it is. The difference is that this attraction is much stronger than most dipole-dipole attractions. It is called hydrogen bonding, and it occurs only in compounds containing hydrogen that is covalently bonded to a small, highly electronegative element (usually nitrogen, oxygen or fluorine).

Hydrogen bonding is an electrostatic attraction between molecules and not a true form of chemical bonding. It is, however, the strongest of the intermolecular forces of attraction. It accounts for some very important physical properties of compounds in which it occurs. Without this strong intermolecular force, water, for example, would be a gas at room temperature, and ice, in all probability, would be more dense than water. You can let your imagination run on the implications for life on our planet, if all water existed in the vapor state and/or ice had a density greater than that of liquid water.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

Name_____

Section_____

INTRODUCTION

Steps 1 through 8 of this investigation refer to a set of demonstrations that your teacher will perform. In order to get the maximum benefit from these demonstrations, you will need to observe the systems very carefully. The main goal of this investigation is to ascertain information on the reactivity of the representative (A-column) elements.

EXPLORATION

1. Your teacher will place sodium metal in water. Observe carefully and record your observations.

2. Was the interaction a physical change or a chemical change? List evidence which supports your choice.

3. Your teacher will perform another procedure using sodium and water. Record your observations.

4. Write a balanced equation for the reaction between sodium and water.

5. How would the reaction rate of lithium in water compare to the reaction rate of sodium in water? Explain your predictions in terms of ionization energy and electronegativity.

6. Carefully describe the behavior of lithium in water, including a balanced equation for the reaction.

7. Predict the behavior of potassium metal in water, based on the information gained from the interactions of lithium and sodium with water.

8. Observations of potassium in water. (Include a balanced equation for the reaction.)

9. Place one piece of aluminum, calcium and magnesium in separate beakers containing 30 to 40 ml of distilled water, allow them to sit for two or three minutes, and record your thorough observations for each. For best comparison try to get the metal pieces about the same size and shape.

10. Heating will sometimes speed up a reaction. If no change (or little change) is noticed after two or three minutes, heat the system and record your observations.

11. Write balanced equations for the chemical reactions observed.

12. Rinse out the beakers used earlier and put 30-40 ml of HCl solution in them. Place the metals in the solutions as before and record your observations.

13. Write a balanced equation for each reaction of the metals used with the acid solution.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

THE IDEA

1. Rank the three I-A metals in order from <u>least</u> reactive to <u>most</u> reactive with water. Offer an explanation for the trend in reactivity. (Laboratory observations will not be sufficient to completely answer this.)

2. How would the behavior of cesium in water compare to that of the other I-A metals used in this activity?

3. Provide reasons for this proposed behavior of cesium.

4. In general, do you think metals are more reactive with water or acid? Why should they be more reactive with one than with the other?

5. How would you expect Li, Na and K to react with an acid solution? Explain your answer.

6. Write the total electron configurations (not electron-dot diagrams) for <u>neutral</u> atoms of Na, Mg and Al.

7. Write the total electron configurations of Na, Mg and Al <u>after</u> they have reacted, including appropriate charges.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

8. What had to happen to Na, Mg and Al to change the configurations between Question 6 and Question 7?

9. List as many relationships as possible between the position on the periodic chart and the reactivity for these metals (I-A, II-A, III-A).

NOTE: It is not necessary to include all six metals in any given comparison, but all six metals should be included in your discussion.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

Name_____

Section_____

EXPANSION OF THE IDEA

HALOGENS AND HALIDES: EXPERIMENT A

1. Test the solubility of NaCl, NaBr, and NaI in water by putting a small amount (enough to cover a match head) of each into a test tube and adding 2-3 ml of water and mixing vigorously. Repeat the test with 2-3 ml of hexane. Record your observations in the table below.

	water	hexane
NaCl	_____	_____
NaBr	_____	_____
NaI	_____	_____

2. Mix the NaCl in water and NaCl in hexane together. Mix the solution vigorously for 15 seconds. Do the same for the NaBr and NaI solutions. Record your observations.

3. Test the solubility of I_2 crystals in water and in hexane in a manner similar to Step 1. Add a few ml of hexane to the water solution of I_2 and mix vigorously. Record your observations.

4. Obtain water solutions of I_2, Br_2 and Cl_2. Place 5 drops of each in separate 13 x 100 mm test tubes and add 1 ml of hexane. Mix vigorously. Repeat the test using 1 ml of I_2, Br_2 and Cl_2 instead of 5 drops. Record your observations in the table below.

NOTE: These solutions should be retained as standards for comparison throughout the experiment. The test tube containing the color standards should, therefore, remain stoppered. If the color in the hexane layer should fade, add another 1 ml of the appropriate halogen water solution and shake.

	5 drops	1 ml
I_2	_____	_____
Br_2	_____	_____
Cl_2	_____	_____

QUESTIONS

1. Classify, in the table below, water, hexane, fluorine (F_2), chlorine (Cl_2), bromine (Br_2), iodine (I_2), sodium fluoride (NaF), sodium chloride (NaCl), sodium bromide (NaBr), and sodium iodide (NaI), by listing them in the appropriate column.

Polar	Non-Polar

2. Use the test tube diagram to illustrate what you would see if equal volumes of hexane and water were poured into a test tube. Label your diagram.

3. Diagram and label what you would expect to find in the test tube of Question 2 if you added Br_2 to the tube or if you added NaBr to the tube.

4. Explain the chemical difference between the chlorine atom of NaCl and the chlorine atom in Cl_2.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

HALOGENS AND HALIDES: EXPERIMENT B

You will first need to make color standards to use for comparison with the reaction test tubes. According to the solubility rules, non-polar substances should dissolve in non-polar substances, so the color obtained in the hexane layer of these color standards should represent the color of the respective elemental halogens. Keep this fact in mind while performing the tests later.

Color Standards

1. Clean and dry three 13 x 100 mm test tubes. Label one test tube chlorine, a second test tube bromine, and a third test tube iodine.

2. Put approximately 1 ml of hexane in each of the test tubes.

3. To the chlorine test tube add approximately 2 ml of chlorine water; to the bromine test tube add approximately 2 ml of bromine water; and to the iodine test tube add approximately 2 ml of iodine water.

4. Stopper each test tube and shake vigorously.

NOTE: These standards should be retained for comparison throughout the experiment. The test tubes containing the color standards should, therefore, remain stoppered. If the color in the hexane layer should fade, add another 1 ml of the appropriate halogen water solution and shake.

5. Put approximately 1 ml of NaF solution in a clean, dry 13x100 mm test tube, 1 ml of NaBr solution in a second test tube, and 1 ml of NaI solution in a third test tube. To each of these test tubes add approximately 1 ml of hexane. Stopper, shake vigorously, and record your observations.

6. Add approximately 1 ml of Cl_2 water to each test tube. Stopper and shake vigorously. Record your observations of the systems after shaking.

7. In which test tubes did chemical changes occur? Support your answers with arguments based on your observations.

8. Write balanced equations for the chemical reactions. Include only the reactants and products in the equations, not the solution media.

NOTE: Dispose of the contents of these test tubes according to your teacher's instructions. Wash the test tubes.

9. Obtain approximately 1 ml of NaCl and NaI solutions in separate test tubes. Add 1 ml of hexane to each test tube, stopper, shake vigorously, and record your observations.

10. Add 1 ml of Br_2 water to each test tube, stopper, shake vigorously, and record your observations after shaking.

11. In which test tubes did chemical reactions occur? Support your answers with arguments based on your observations.

12. Write balanced equations for the chemical reactions. Include only the reactants and products, not the solution media.

NOTE: Dispose of the contents of these test tubes according to your teacher's instructions. Wash the test tubes.

13. Obtain approximately 1 ml of NaF, NaCl and NaBr solutions in separate test tubes. Add 1 ml of hexane to each test tube, stopper, shake vigorously, and record your observations.

14. Add 1 ml of I_2 water to each test tube, stopper, shake vigorously, and record your observations after shaking.

15. In which test tubes did chemical reactions occur? Support your answers with arguments based on your observations.

16. Write balanced equations for the chemical reactions. Include only the reactants and products, not the solution media.

NOTE: Dispose of the contents of the test tubes according to your teacher's instructions. Wash out the test tubes containing the color standards and the reaction test tubes.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

QUESTIONS

1. Write total electron configurations (not electron-dot diagrams) for neutral F, Cl, Br, and I atoms.

2. When halogen atoms react chemically to form ions, what changes occur in their electron configurations?

3. Rank the halogens (bromine, chlorine and iodine) in order from <u>least to most</u> reactive. Offer a thorough explanation for this trend in reactivity. (Laboratory observations will not be sufficient to completely answer this.)

4. Predict the reactivity of fluorine compared to the other members of the VII-A family. Offer reasons for your proposed behavior of fluorine.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

EXPANSION OF THE IDEA

READING a: REACTIVITY OF METALS

In the Reading Active and Noble Metals, some aspects of the comparative chemical reactivity of several metals were discussed. Those metals, in order from most reactive to least reactive, were:

zinc
iron
nickel
tin
lead
copper
mercury
silver
gold
platinum

Notice that all of these metals are included in the groups of transition metals. They are metals that have practical uses, either in their pure elemental states or in solid solutions with other elements (alloys). This is generally true of not only these ten but most transition metals for two important reasons: (1) they tend to have good structural characteristics (that is, they are hard and strong); and (2) they have relatively low chemical reactivities in the situations where they are used.

If it were not for the relatively low reactivity of these metals, their use would be much more restricted, especially when such uses exposed them to reactants such as oxygen, water vapor, and pollutant acids. As it is, these metals generally have to be protected in some way to limit the rate of chemical reactions that would cause a loss of the desirable characteristics for which the metals are employed. Paints, varnishes and other coatings are sometimes used for this protection, but some metals enhance this protection process by the chemical formation of a tightly-adhering oxide film that inhibits further reaction. This often happens with aluminum and iron cookware--a protective situation that can be destroyed by cleaning with strong solvents or abrasive substances.

In order to judge the comparative chemical reactivity of metals it will first be helpful to review how metals characteristically react. Metals, as you recall, tend to form positive ions, and this positive ion formation would, of course, involve the loss of electrons. In fact, you will notice when you look back at your common cation sheet, that all the ions listed there (except ammonium--NH_4^+) are metals! Logically, therefore, the easier it is to pull

electrons off a metal atom, the more reactive that metal would be. Stated in other words, the more readily a metal gives up electrons, the more reactive it is. Given the way that metals typically react, it is easy to see why they would have to be at the lower end of the electronegativity scale. In fact, metals are sometimes said to be electropositive, with the more active metals being more electropositive and the more noble metals being less electropositive.

There are two broad factors that should affect the tendency of an atom to give up electrons:

(1) The size of the atom, and
(2) The number of valence electrons.

The size of the atom will generally vary inversely as its tendency to lose electrons. In other words, the larger the atom, the easier it should be to remove an electron--all other things being held constant. The trends in ionization energies from top to bottom in a family support this idea. Conversely, the smaller atom should hold on to its electrons more strongly. Regarding the number of valence electrons, we would logically conclude that one metallic element would generally be more reactive than another metallic element, if the first one had fewer valence electrons than the second.

The metals we have used in Investigtion 19 are all more reactive than any of the metals listed in the Reading referred to earlier. It can be noted that most of these metals (aluminum is a notable exception) are not very useful in their metallic states because of their high reactivities. The I-A metals must even be kept under kerosene in order to prevent rather rapid reaction with the air. Let's compare our laboratory observations to the factors proposed in this reading for determining metal reactivity. The rate of reaction of the metals sodium, magnesium and aluminum decreased in the order that these elements are listed. Notice that sodium has one valence electron, magnesium has two, and aluminum has three. It appears that the number of valence electrons factor is significant in judging reactivity. This can also be seen by comparing the reaction rate of calcium (two valence electrons) with potassium (one valence electron). The other factor--atomic size--and its relationship to reactivity can be judged by comparing the reactions of lithium, sodium and potassium, magnesium and calcium. In these examples the smaller atoms were the least reactive.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

If you look at the list of metals from ACTIVE AND NOBLE METALS again and then find their positions on the periodic chart, you will see that the trends in metal reactivities that have been implied in this reading apply mainly to A-column metallic elements. However, definite reactivity trends for these A-column elements can be quite useful, and they are worthy of remembering.

INVESTIGATION 19: PERIODICITY AND REACTIVITY

EXPANSION OF THE IDEA

READING b: REACTIVITY OF NON-METALS

During Investigation 5 you performed a burning reaction between magnesium and air. The speed of that reaction was controlled by limiting the exposure of air to the magnesium. If you lifted the crucible cover off your system while it was still very hot, you noticed that the remaining magnesium burst into flames, giving off large quantities of white smoke. Even if you waited to lift the lid until the system was cool enough not to burn, you noticed that white smoke (actually small particles of MgO) had deposited on the inside of the crucible cover. Both reactions--the one performed in the closed system and the one performed in the open air--were combustion reactions, but the products of these reactions differed. While the burning in open air appeared to produce only magnesium oxide, the reaction in the closed system resulted in the production of both magnesium oxide and magnesium nitride! Clearly, the oxygen and nitrogen both came from the air which contains them, but why did nitrogen react in one system and not in the other? In an atmosphere that contains almost 80% nitrogen and only about 20% oxygen, why should magnesium oxide be the predominant product?

The answers to both of these questions are related. Since the reaction with oxygen is preferred when there is plenty of oxygen available, we could assume that the activation energy required for that reaction is lower than that required for the nitrogen reaction. In support of this idea is the observation of the formation of some magnesium nitride in an oxygen starved system, when the system receives continued strong heating. It appears that oxygen will react more readily (under milder reaction conditions) than nitrogen--at least it will when the other reactant is magnesium. We find, in fact, that this situation is generally true; that is, oxygen is more chemically active than nitrogen. That is not to say that nitrogen is unreactive. Nitrogen is an element contained in thousands of different compounds. Ammonium nitrate (NH_4NO_3), for example, contains two ions, and each of those ions contains nitrogen. All amino acids--the building blocks of protein molecules--contain nitrogen, and a list of nitrogen-containing compounds could go on and on. The point is simply that oxygen will react more readily than nitrogen.

If we consider only the formation of the nitride and oxide ions in reactions of these elements, it is easy to see why oxygen would be more reactive than nitrogen. Nitrogen needs to gain three electrons, while oxygen needs to gain only two. Furthermore, the additional proton of the oxygen nucleus should enhance its ability to attract electrons from

other elements. If we continued this line of reasoning to the VII-A element of the second period (fluorine), we would have to predict that fluorine is more reactive than either nitrogen or oxygen. This conclusion would result from the magnitude of the nuclear charge (9 for fluorine) and the fact that only one additional electron is required for fluorine to achieve a noble gas electron configuration.

Is fluorine more reactive than nitrogen and oxygen? The answer is an emphatic yes! Fluorine is, in fact, the most reactive of all the chemical elements. More different elements will combine with fluorine than with any other element. Our laboratory data clearly illustrate the high reactivity of fluorine compared to the other VII-A elements.

It appears that the reactivity of non-metals will be related to the same two factors that the reactivity of metals was:

(1) number of valence electrons, and
(2) atomic size.

We have seen that metals and non-metals are consistently opposite of each other--physically and chemically. If that pattern of oppositeness is consistent, we would expect the effects of these two factors to have an opposite effect on reactivity in metals and non-metals. This is, in fact, the situation. V-A elements are less reactive than VI-A elements, and VI-A elements are less reactive than VII-A elements. They have five, six and seven valence electrons, respectively, so we see that reactivity increases (in a period of non-metals) as the number of valence electrons increases. We have seen a very nice pattern of decreasing reactivity in the VII-A elements, as the size of the atom increases; that is, iodine is less reactive than bromine, and so forth. This fits well with the idea of a decrease in attraction between opposite charges as the distance increases. The less the attraction, the lower the reactivity of the non-metal. A general trend in reactivity of non-metals arises as a result of the foregoing considerations. The farther up and to the right a non-metal is on the periodic table, the more reactive it is. "Aha!", you say. "What about the noble gases?" Welllll. . . .